Contents

Chapter 1

States of Matter and Phase Changes

Student Learning Outcomes

After studying this chapter, students will be able to:

- Define chemistry as the study of matter, its properties, composition, and interactions with other matter and energy. Or Study of earth (solids), Air (gases), Sea (liquids) and Sky (plasma) and their interaction with each other.
- Explain with examples that chemistry has many sub-fields and interdisciplinary fields. Some examples include:
  - Biochemistry
  - Medicinal Chemistry
  - Polymer Chemistry
  - Geochemistry
  - Environmental Chemistry
  - Analytical Chemistry
  - Physical Chemistry
  - Organic Chemistry
  - Inorganic Chemistry
  - Nuclear Chemistry
  - Astrochemistry
- Define matter as a substance having mass and occupying space.
- State the distinguishing macroscopic properties of commonly observed states of solids, liquids and gases in particular density, compressibility and fluidity.
- Identify that state is a distinct form of matter (examples could include familiarity with plasma, intermediate states and exotic states e .g., BEC or liquid crystals)
- Explain the allotropic forms of solids (some examples may include diamond, graphite, and fullerenes)
- Explain the difference between elements, compounds and mixtures
- Identify solutions, colloids, and suspensions as mixtures and give an example of each
- Explain the effect of temperature on solubility and formation of unsaturated and saturated solutions

1.1 What is Chemistry?

Chemistry is the brance of science which deals with the properties, composition and the structure of substances. It also deals with the physical and chemical changes in matter and the laws or principles which govern these changes.

Determination of composition represents finding out percentages of elements and compounds in a sample of matter. Structure of matter means the arrangement of atoms in matter. Both physical and chemical changes may be brought about by absorption or evolution of energy.

Branches of Chemistry

To understand the widely spread complex subject of chemistry and to concentrate on its specific aspects, chemistry is divided into many distinct branches. These branches have distinct areas of study for the scientists to focus on and to achieve breakthroughs and advancements.

1. Physical Chemistry

This branch investigates how substances behave at atomic and molecular levels. It provides clear explanation as to how fundamental physical laws governing our world cause atoms and molecules to show specific characteristics and in turn react to give huge structures related to life. Physical chemistry is also used to predict and change the rates of reactions and thus optimize the conditions to carry out the reactions on industrial scale.

2. Inorganic Chemistry

It is the study of the synthesis, composition, properties and structure of elements and compounds that contain little or no carbon. An inorganic substance can be composed of metals, nonmetals or a mixture of these, salts, acids and bases. Inorganic compounds are used as fertilizers, medicines, catalysts, pigments, coatings and much more.

3. Organic Chemistry

It is the branch of chemistry that deals with the carbon compounds (hydrocarbons and their derivatives) other than its simple salts like carbonates, bicarbonates, oxides and carbides. In this branch, we study the structure, formation, properties, composition and reactions of carbon containing compounds. Organic compounds are found in all forms of life and are also essential for life.

4. Environmental Chemistry

It is the scientific study of the chemical and biochemical phenomena that occur in this planet. In this subject, we study the sources, reactions, effects and fates of chemical species in the air, soil and water environments. Without this, it would be impossible to study the effects that humans have on the environment through the release of chemicals. It helps in understanding the causes, effects and solutions of different types of pollution.

5. Analytical Chemistry

This branch of chemistry deals with the analysis of different substances. It involves separation, identification and determination of the concentration of the components present in material things. Nowadays the field of analytical chemistry generally involves the use of modern and sophisticated instruments to analyze the matter.

6. Biochemistry

It is the branch of chemistry in which we understand life through chemical processes. It is the study of chemical substances and vital processes occurring in living organisms. Biochemistry provides insights into the structure and function of molecules such as proteins, carbohydrates, lipids and nucleic acids.

7. Nuclear Chemistry

Nuclear chemistry deals with the reactions taking place in the nucleus of an atom. It deals with radioactivity, nuclear processes and transformation in the nuclei of atoms. Nuclear chemistry has many applications in agriculture, medicine, industry and research.

8. Polymer Chemistry

Polymers are large molecules made by linking together a series of building blocks. Polymer chemistry focuses on the properties, structure and synthesis of polymers and macromolecules. Many materials present in the living organisms including proteins, cellulose and nucleic acids are naturally occurring polymers.

9. Geochemistry

The study of chemical composition of Earth and its sources and minerals is called geochemistry. Apart from its use in minerals exploration, geochemical

mapping today has applications in environmental monitoring, forestry and medical research.

Interesting Information!

Geothermal heat pump uses a pump to transfer underground water into the buildings during the winter to heat them and in the summer to cool them.

10. Medicinal Chemistry

In this branch of chemistry, the chemist tries to design and synthesize medicines or drugs which are beneficial for mankind. It includes the discovery, delivery, absorption and metabolism of drugs in human body.

11. Astrochemistry

It is the study of molecules and ions recurring in space and interstellar space. In this discipline we study the abundance and reactions of atoms, molecules and ions in the universe and interaction of these species with radiation.

Exercise

A lunar mission has recently brought samples from the Moon. The following experiments were then carried out on it. Point out the branch of chemistry these experiments are related to.

Experiment	Branch of Chemistry
1. Determining its composition	
2. Studying the physical properties of materials it contains	
3. Carrying out chemical reactions with usual inorganic reagents	

1.2 States of Matter

This world is made up of matter and energy. Energy is non-material in nature. Anything other than energy which carries weight and occupies volume is called matter. We encounter material things everywhere in all sorts of different and distinct forms.

A state of matter is one of the many distinct forms in which matter can exist. We observe four states of matter in everyday life: solid, liquid, gas and plasma. Apart from these, there are more states of matter which we do not see in our everyday life.

The three primary states of matter are the solid, liquid and gaseous. They are different from each other due to different strength of intermolecular forces, the arrangement of particles and the distance between the particles Fig (1.1). In gases, molecules are very widely apart with no order whatsoever and very weak intermolecular forces. All these features make gases easily compressible. Their densities are obviously very low.

The liquids, on the other hand, have molecules which are closely attached but moving randomly. There exists significant intermolecular forces between their molecules. Liquids are therefore not easily compressible and their densities have higher values than those of gases.

Solids have a definite shape and a fixed volume. Particles in solids are closely packed and have very strong interatomic or intermolecular attractions. The particles in solids remain fixed at their positions where they can oscillate about their mean positions. Solids are relatively incompressible and rigid.

The densities of solids are very high. Solids are the only state of matter which do not need any container to be stored. In crystalline solids, particles are perfectly arranged and strongly bonded. This makes them almost incompressible. They are the most dense substances.

Plasma is not so generally seen form of matter. It is composed of particles with very high kinetic energy. It exists in fluorescent tubes, lightning and welding arcs. Plasma can be considered as a partially ionized gas containing electrons, ions, photons, etc.

Matter also exits in intermediate states where liquid meets gas and liquid meets solid, for example supercritical fluids, liquid crystals and graphene. Supercritical fluids are highly compressed states which show both properties of gases and liquids. Chemical reactions which may not be carried out in conventional solvents, may possibly be carried out in supercritical carbon dioxide.

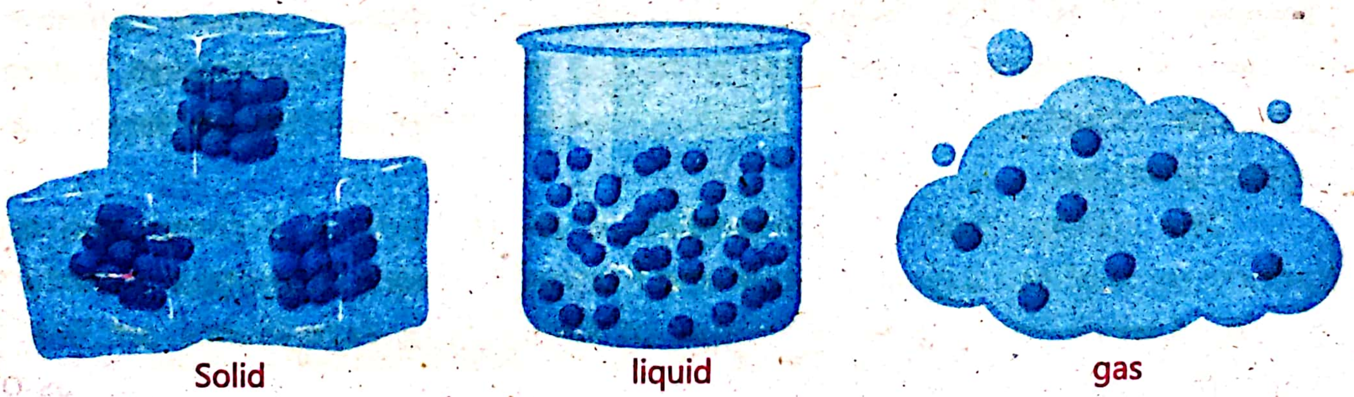


Fig (1.1) States of Matter

Liquid crystal is a state of matter whose properties are between those of conventional liquids and those of crystalline solids. Liquid crystals are used in display devices including computer monitors, clocks, watches and navigation systems. Graphene is an example of two-dimensional crystal, a single layer of carbon atoms arranged in a hexagonal pattern. Graphene is a tough, flexible and light material with a high resistance Fig (1.5).

States of matter that are not commonly encountered are called exotic states of matter. Examples are dark matter, Bose-Einstein condensate, nuclear matter, quantum spin liquid and many others.

1.3 Element, Compound and Mixture

Matter in this world exists in the form of elements, compounds and mixtures. Element is the simplest form of matter. It is a pure substance containing the same kind of atoms. It cannot be broken down into simpler substances by ordinary chemical reactions. Elements exist in all three forms; solid, liquid and gas. Most of the elements found in this world exist in solid form. Liquid and gaseous elements are very few in number as compared to solids, Fig (1.2).

Elements may be a metal, a non-metal, a metalloid and a noble gas. Elements can also exist in the form of atoms, molecules, ions and isotopes. Examples of important elements are sodium, potassium, magnesium, calcium, carbon, silicon, nitrogen, oxygen, chlorine, helium, copper, gold, zinc, silver, nickel, cobalt, mercury, bromine and iodine, etc.

Compound is also a pure substance. It is made up of two or more different chemically elements combined in a fixed ratio. When elements come together, they react with each other and form chemical bonds that are not easy to break, Fig (1.3).

Compounds may be molecular, ionic, intermetallic and coordination complexes. Compounds may also be inorganic and organic in nature. Examples of important compounds are water, ammonia, methane, carbon dioxide, sodium carbonate, potassium chloride, starch, proteins, carbohydrates, mineral acids, organic acids, etc.

The composition and properties of an element or a compound are uniform throughout a given sample and from one sample to another. A mixture is formed when more than one types of elements or compounds are mixed together in any ratio. Air, soil, milk and tap water are everyday examples of mixtures. A mixture may be homogeneous or heterogeneous. A solution of salt

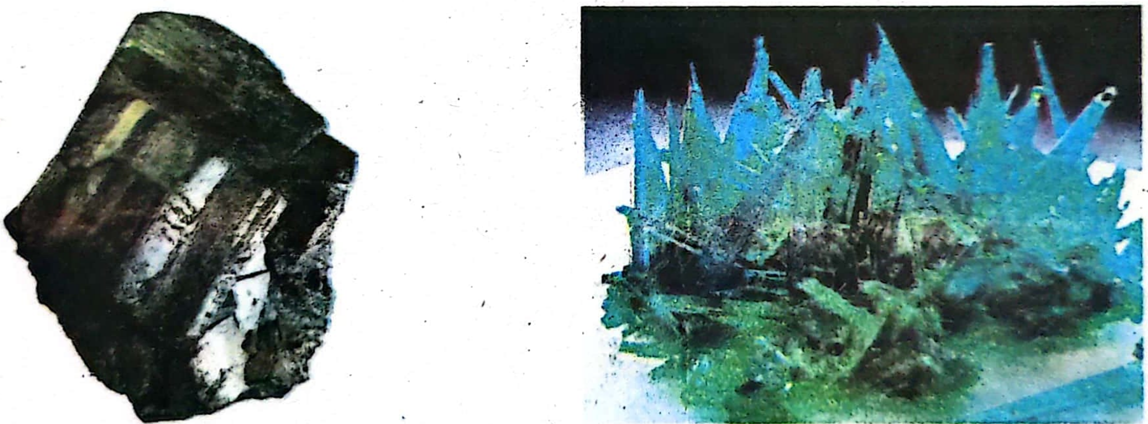
and water is an example of homogeneous mixture because its concentration is uniform throughout. A sample of rock is an example of heterogeneous mixture because the concentration of its constituents is different in its different parts. Rocks are composed of different types of minerals such as granite, mica and limestone, Fig (1.4).

Interesting Information!

Many elements are found in nature but some are artificial. Technetium was first element created by scientists in the laboratory.

Iron rods | Copper wire | Screw and nuts of zinc

Fig (1.2): Elements



Iron sulphide | Copper chloride

Fig (1.3): Compounds

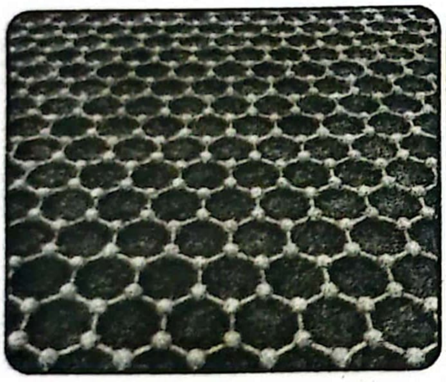
Rock | Chocolate

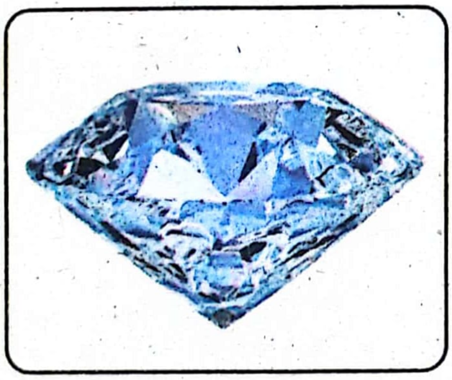
Fig (1.4): Mixtures

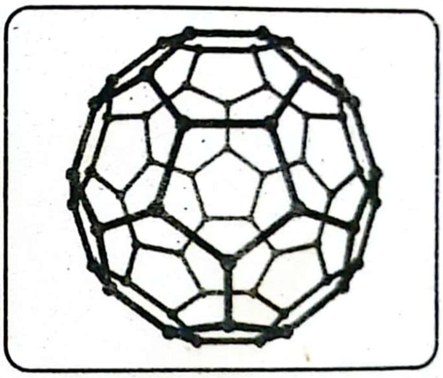
1.4 Allotropic Forms of Substances

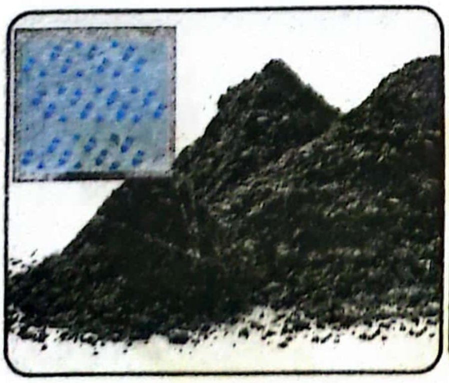
Elements may exist in more than one structural forms which can exhibit quite different physical and chemical properties. These forms are called allotropic forms and phenomenon is called allotropy. Element oxygen exists in two allotropic forms namely oxygen (O_2) and ozone (O_3).

Similarly, carbon exists in three main allotropic forms, diamond, graphite and Buckminster fullerene. Diamond has a giant macromolecular structure whereas graphite has a layered structure of hexagonal rings of carbon. Buckminster fullerene (C_{60}) consists of spheres made of atoms arranged in pentagons and hexagons. Fullerenes are stable at high temperatures and high pressures. Being covalent in nature, they are soluble in organic solvents. The fullerene structure is unique in that the molecule is not charged, has no boundaries and has no unpaired electrons. They have a cage like structure. Fullerene C_{60} has a low melting point. It is soft and cannot conduct electricity. Element sulphur also exists in two crystalline allotropic forms i.e. rhombic and monocline; the former is more stable than the latter, Fig (1.5).


Graphene


Diamond


Buckminsterfullerene


Graphite


Rhombic Sulphur

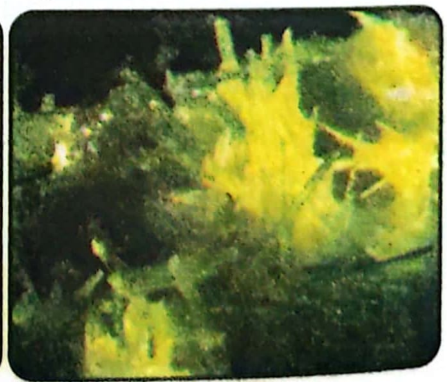

Monoclimic Sulphur

Fig (1.5) Allotropic Forms of Carbon and sulphur

1.5 Differences between Elements, Compounds and Mixtures

Elements	Compounds	Mixtures
An element is the simplest form of matter. It is a pure substance containing the same kind of atoms.	A compound is a pure substance. It is formed by the chemical combination of two or more atoms of different elements.	Mixture is an impure compound. A sample of matter having more than one type of elements or compounds mixed together in any ratio, is called a mixture.
It is not possible to break down an element into simpler particles by ordinary chemical reactions.	In a compound, the atoms of elements must combine together by a fixed ratio by weight. For example, in water (H_2O) hydrogen and oxygen are present in a fixed ratio of 1:8 by weight.	Each component of a mixture retains its identity and specific properties.
When an element exists in the form of atoms, it is represented by a symbol. For example, sodium and calcium are represented by their symbols Na and Ca.	It is possible to break a compound into its constituent elements by a chemical reaction. For example, ammonia can be converted back to nitrogen and hydrogen by a suitable chemical reaction.	A mixture may be homogeneous or heterogeneous. For example, the solution of common salt in water is a homogeneous mixture while a sample of rock is a heterogeneous mixture.
Gaseous elements exist in the form of independent molecules, for example, nitrogen (N_2), oxygen (O_2) and chlorine (Cl_2). Noble gases, however, exist as mono atomic molecules.	The properties of a compound are always different from the elements from which it is formed. For example, the properties of water are different from those of hydrogen and oxygen.	The components of a mixture are not chemically bound together and they can be separated by physical methods.

For example, helium (He) and argon (Ar).	Compounds exist in the form of molecules, for example hydrogen chloride (HCl), ammonia (NH_3) and water (H_2O). Compounds may also exist as network arrangement of their atoms. For example, ionic compounds like NaCl and covalent compounds like sand (SiO_2).	The properties of a mixture are the sum of those of its components.

Exercise
1. Which elements are found in pure state on Earth?
2. Which elements are present in very small amounts on Earth?

1.6 Solution, Colloidal Solution and Suspension

A solution is such a mixture in which solute particles are completely homogenized in the solvent e.g. dissolution of sodium chloride or copper sulphate in water. The solute particles in such a solution cannot be seen by the naked eye. If solution is filtered, these particles pass through the pores of filter paper leaving no residue. Such a solution is called a true solution. A suspension, on the other hand, is a mixture in which solute particles do not dissolve in the solvent. We can actually see these particles. If a suspension is kept for some time, these particles settle down. Again, if this suspension is filtered, the particles in it cannot pass through the pores of filter paper and can be collected as a residue. A mixture of chalk in water is an example of suspension. Fig (1.6).

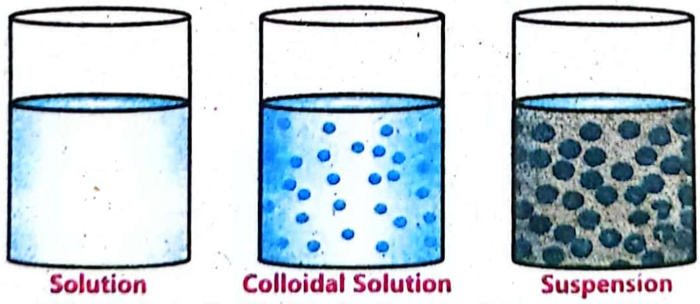


Fig (1.6) Colloidal Solution and Suspension

Besides a true solution and a suspension, there is a third form of solution called colloidal solution. In this type, the solute particles do not homogenize with solvent. These particles are a little bit bigger than the solute particles present in the true solution but not big enough to be seen with a naked eye like the particles present in a suspension. If kept for some time, the particles of a colloidal solution do not settle down. On filtration, these particles pass through the filter paper like

the particles of a true solution. Starch solution and white of an egg are the common examples of colloidal solutions.

1.7 Formation of Unsaturated and Saturated Solutions

Take about 100 g of water in a beaker, add to it 5g of table sugar, and stir it. The sugar will dissolve in water. Then add another 5g of sugar and stir. This will also dissolve. This solution is called an unsaturated solution. A solution which can dissolve more amount of a solute at a particular temperature is called an **unsaturated solution.** Continue adding sugar in the above solution. As the quantity of sugar in water increases, its dissolution will become more difficult. A stage comes when no more sugar will dissolve in water at this temperature. Any more sugar added at this stage will settle down at the bottom of the beaker. This solution is called a saturated solution at a particular temperature of the process.

A solution in which the maximum amount of the solute has been dissolved in a particular amount of a solvent at a particular temperature is called a **saturated solution.**

Different solutes have different solubilities in a particular solvent. For example, if the solutions of table sugar and sodium chloride are prepared, it is found that 36 g of sodium chloride can dissolve in 100 g of water at 20°C to give a saturated solution. Against this 203.9 g of table sugar can dissolve in water at 20°C to give its saturated solution. In other words, the solubility of table sugar in water is far greater than that of sodium chloride at 20°C. This is due to the fact that sugar molecules are larger than salt ions, so more water molecules can surround a single sugar molecule, causing it to dissolve in larger amounts.

Interesting Information!

Mixtures are closely related to our everyday lives. The air we breathe, the foods we generally consume, the fluids in our body, the solids like steel we use, are all either homogeneous or heterogenous mixtures.

1.8 Effect of Temperature on the Solubility of Solutes

The solubility of a solute is the amount of solute which can dissolve in 100g of a solvent at a particular temperature.

Change in temperature has different effects on the solubility of different compounds. Usually the solubility increases with the increase in temperature but it cannot be taken as a general rule. There are a large number of compounds whose solubility in water increases with the increase in temperature e.g.

potassium nitrate (KNO_3), silver nitrate ($AgNO_3$) and potassium chloride (KCl) etc. The solubility of sodium chloride in water does not increase appreciably with the increase in temperature. The solubility of compounds like lithium carbonate (Li_2CO_3) and calcium chromate ($CaCrO_4$) decreases with the increase in temperature. The solubility of gases in water also decreases with the increase in temperature, Fig (1.6).

Similarly, the solubilities of copper sulphate and sodium nitrate also increase with increase in the temperature. However, the solubility of calcium hydroxide decreases with the increase in temperature.

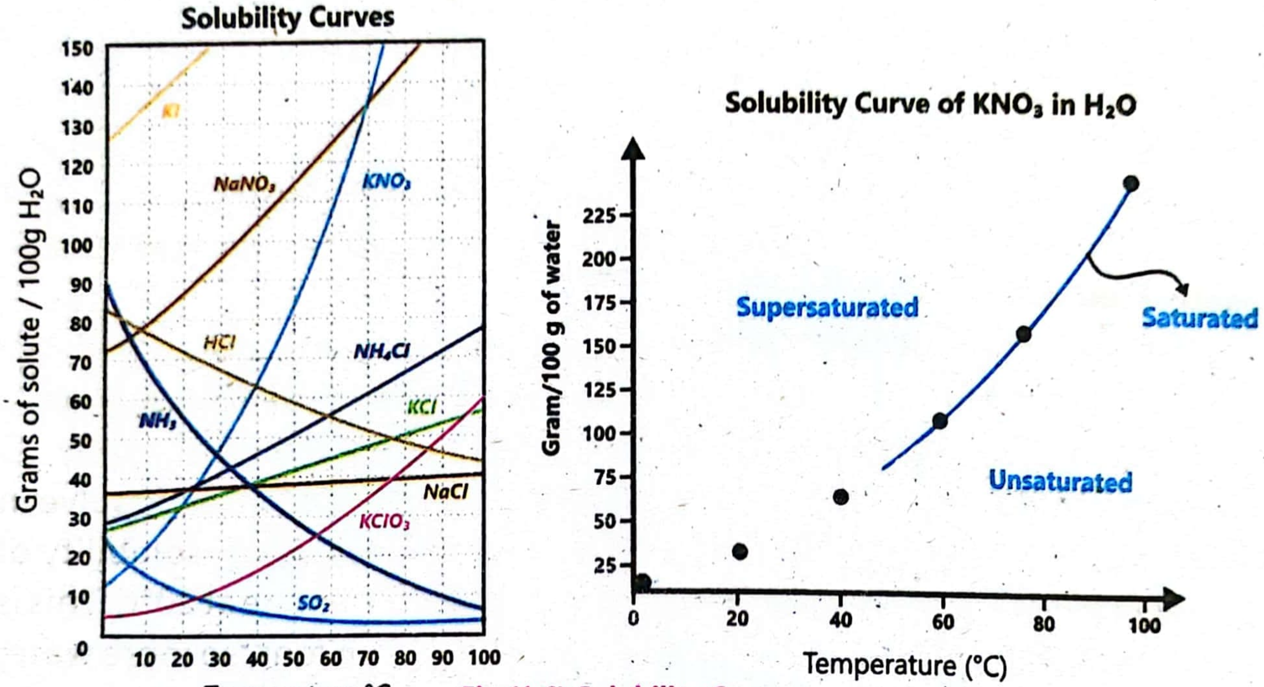
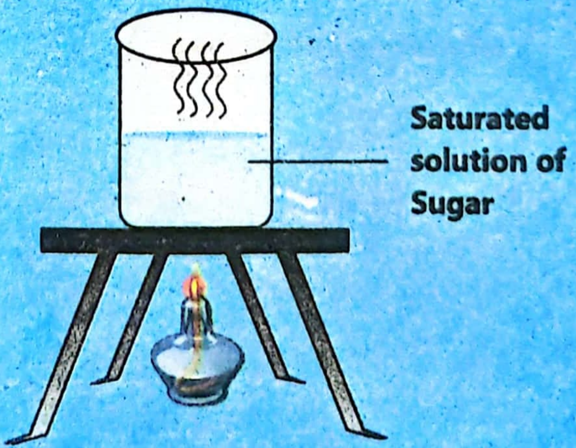


Fig (1.6) Solubility Curves

Activity

Take 100 g of water in a beaker and prepare saturated solution of sugar at room temperature. Heat the beaker on a spirit lamp. Add a little more sugar in it and stir it. Will this sugar be dissolved in it? You will notice that by heating the solution the quantity of sugar dissolved in water has increased i.e. the solubility of sugar has increased.

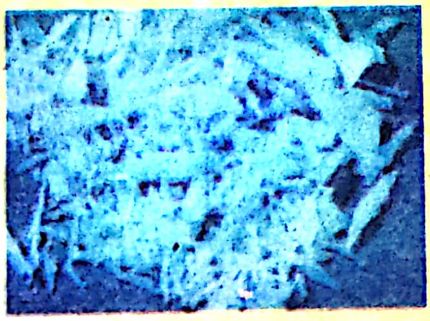


Exercise

How variation of solubility at different temperatures can be useful for us?

Interesting Information!

1. The increase in the solubility of solids in liquids with increase in temperature may be used to purify them. Pure solids commonly appear as beautifully shaped crystals.



Crystals of Potassium Nitrate

2. Generally the solubility of gases decreases with increase in temperature. Carbon dioxide gas is more soluble in water at low temperature. Soda water bottles are thus stored in the refrigerator to keep carbon dioxide gas dissolved in water for a longer period of time.

Key Points

1. Chemistry is that branch of science which deals with the composition of matter, changes in matter and the laws which govern these changes.
2. To understand the vast and complex subject of Chemistry, it is divided into many branches. Physical chemistry, inorganic chemistry and organic chemistry are its main branches among so many others.
3. Matter exists mainly in three states: solid, liquid and gas. They are different from each other due to different distances between the particles which they contain.
4. Plasma is regarded as the fourth state of matter which is not normally observed in this world. Most of the matter present in the rest of the universe exists in this state.
5. Matter also exists in the intermediete states which are at the borderline of its two principal states, between liquid and gas or between liquid and solid. Supercritical fluids and liquid crystals are some examples of such states.
6. Matter exists in the form of distinct entites called elements, compounds and mixtures.
7. Elements, compounds and mixtures have distinct properties individually and they are very different form one another.
8. Solutions, suspensions and colloidal solutions are different forms in which the mixtures usually exist. They have their own characteristic properties.
9. Different solutes have different solubilities in a particular solvent.
10. The solubility of a solute is the amount of solute which can dissolve in 100g of a solvent at a particular temperature.
11. Change in temperature has different effects on the solubility of different compounds.

Exercise

1. Tick (✓) the correct answer.

(i) Matter is present in neon signs in the state of:

(a) Supercritical fluid (b) Plasma

(c) Gas (d) Liquid crystal

(ii) Hazardous effects of shopping bags are studied in:

(a) Geochemistry (b) Inorganic Chemistry

(c) Analytical Chemistry (d) Environmental Chemistry

(iii) The man-made polymer is:

(a) Starch (b) Polystyrene

(c) Protein (d) Cellulose

(iv) The crystals of which substance has rhombic shape?

(a) Brass (b) Sulphur

(c) Graphite (d) Bronze

(v) Which liquid among the following is a colloidal solution?

(a) Milk (b) Slaked lime used for white wash

(c) Vinegar solution (d) Mixture of AgCl in water

(vi) Which of the following is a heterogeneous mixture?

(a) A solution of calcium hydroxide in water

(b) A solution of potassium nitrate in water

(c) Hot chocolate

(d) Concrete mixture

(vii) A state of matter whose properties are between those of liquids and crystalline solids:

(a) Liquid crystal (b) Supercritical fluid

(c) Plasma (d) Dark matter

(viii) When the tiny visible particles of a substance are dispersed through a medium, the mixture is named as:

(a) True solution (b) Colloid

(c) Suspension (d) Saturated solution

(ix) A solution of $KClO_3$ has a solubility of about 13.2g per 100 cm^3 at 40°C. How its solubility will be affected, if you decrease the temperature?

(a) The solubility will increase

(b) The solubility will decrease

(c) The solubility will remain the same

(d) The solubility will first increase with temperature and then it will decrease

(x) You are studying the rate of hydrolysis of starch under different conditions of temperature. In which branch of chemistry this topic will fall?

(a) Organic Chemistry (b) Analytical Chemistry

(c) Biochemistry (d) Physical Chemistry

2. Questions for Short Answers

i. Why is there a need to divide Chemistry into many branches? Give three reasons.

ii. Reactions may take place due to electrons present outside the nucleus or they may take place inside the nucleus. Which branches of Chemistry cover these two types of reactions?

iii. What types of problems are solved in analytical chemistry?

iv. Both graphite and graphene have hexagonal layered structures. What is the difference?

v. Why are supercritical fluids important?

vi. In which state does matter exist in the Sun?

vii. What is the importance of graphene?

viii. Which form of matter do most of the material things in this world belong to?

3. Constructed Response Questions

i. How does a supercritical state look like?

ii. In what way is plasma created in a fluorescent tube?

iii. Most of the molecules we study in biochemistry are organic in nature. Where does the difference exist in organic and biochemistry branches of Chemistry?

iv. Give the reason of brilliance shown by diamond. Can you improve it?

v. Explain the dissolution of sodium chloride in water.

vi. Why do different compounds have different solubilities in water at a particular temperature?

vii. Why NaCl can not be crystallized from water just like KNO_3?

viii. Why graphite is slippery to touch? Which property of graphite enables it to be used as lubricant?

4. Descriptive Questions

i. Mention the name of the branch of Chemistry in which you will study each of the following topics.

(a) Rate of a reaction
(b) Digestion of food in human body
(c) Properties of plasma
(d) Ecosystem
(e) Reactions taking place during fireworks
(f) Measurement of the absorption of wavelength with the help of ultraviolet spectrometer

ii. What are allotropic forms? Explain the allotropic forms of carbon and sulphur. How does coal differ from diamond?

iii. What are supercritical fluids. How are they different from ordinary liquids?

iv. Define solubility of a solute. How does the solubility of solutes change with the increase in temperature?

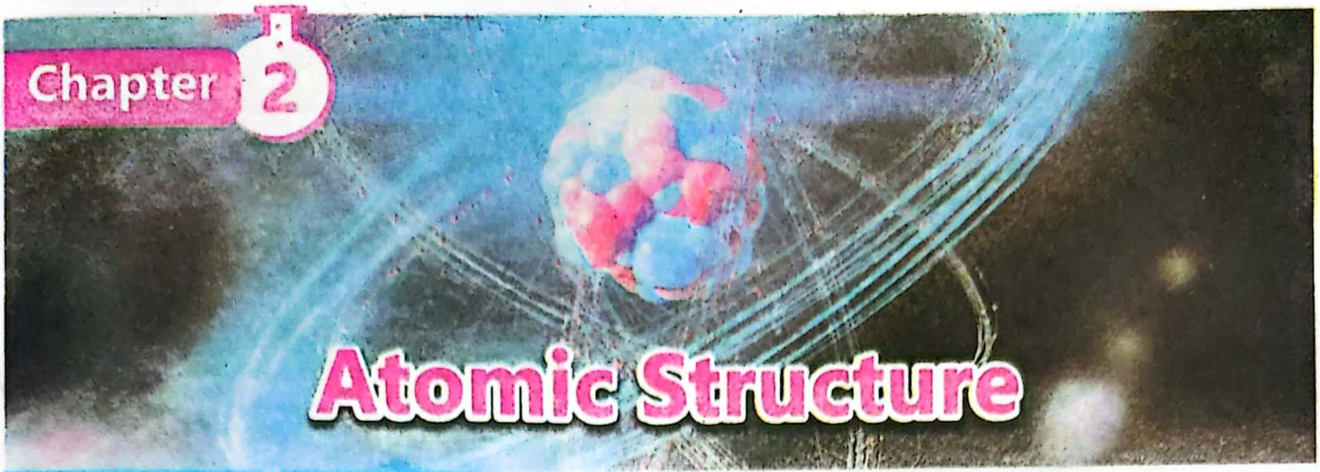
v. What types of movements are present in gaseous and liquid molecules?

vi. Differentiate between the areas which are studied under inorganic and organic chemistry.

5. Investigative Questions

i. Preparation of solutions leads to an important process in chemistry which enables us to purify a compound through crystallization. Describe a process in which potassium nitrate is purified by crystallizing it in water.

ii. Graphene is called a miracle material and it is the material of the future. Which of its many properties makes it very useful in electronics?



Chapter 2

Atomic Structure

Student Learning Outcomes

After studying this chapter, students will be able to:

- Explain the structure of the atom as a central nucleus containing neutrons and protons surrounded by electrons in shells
- State that, orbits (shells) are energy levels of electrons and a larger shell implies higher energy and greater average distance from nucleus
- State that electrons are quantum particles with probabilistic paths whose exact paths and locations cannot be mapped (with reference to the uncertainty principle}
- Explain that a nucleus is made up of protons and neutrons held together by strong nuclear force
- Explain that an atomic model is an aid to understand the structure of an atom.
- State the relative charge and relative masses of a subatomic particles (an electron, proton and neutron)
- Interpret the relationship between a subatomic particle, their mass and charge.
- Illustrate the path that positively and negatively charged particles would take under the influence of a Uniform Electric Field.
- Define proton number / atomic number as the number of protons in the nucleus of an atom.
- Explain that the proton number is unique to each element and used to arrange elements in periodic table
- State that radioactivity can change the proton number and alter an atom's identity
- Define nucleon number / atomic mass as sum of number of protons and neutrons in the nucleus of an atom.
- Define isotopes as different atoms of the same element that have same number of protons but different neutrons
- State that isotopes can affect molecular mass but not chemical properties of an atom
- Determine the number of protons and neutrons of different isotopes
- Define relative atomic mass as the average mass of isotopes of an element compared to 1/12th of mass of an atom of Carbon-12
- State that isotopes can exhibit radioactivity

- Discuss the importance of isotopes using carbon dating and medical imaging as examples. Describe the formation of positive (cation) and negative (anion) ions from atoms.
- Interpret and use the symbols for atoms and ions
- Calculate relative atomic mass of an element from relative masses and abundance of isotopes,
- Calculate the relative mass of an isotope given relative atomic mass and abundance of all stable isotopes.

Elements are very different from one another. A large number of elements exist as solids and very few are present as liquids while the rest exist as gases. All these elements are however made up of atoms.

Have you ever thought why these elements are so different from one another? Sulphur looks very different from gold which, in turn, is very different from bromine (Fig: 2.1). Similarly iron is a heavy metal while aluminium and zinc are light metals. Metals are mostly lustrous while non-metals like sulphur and carbon appear dull. The difference in the properties of elements is due to the difference in the properties of their constituent atoms.

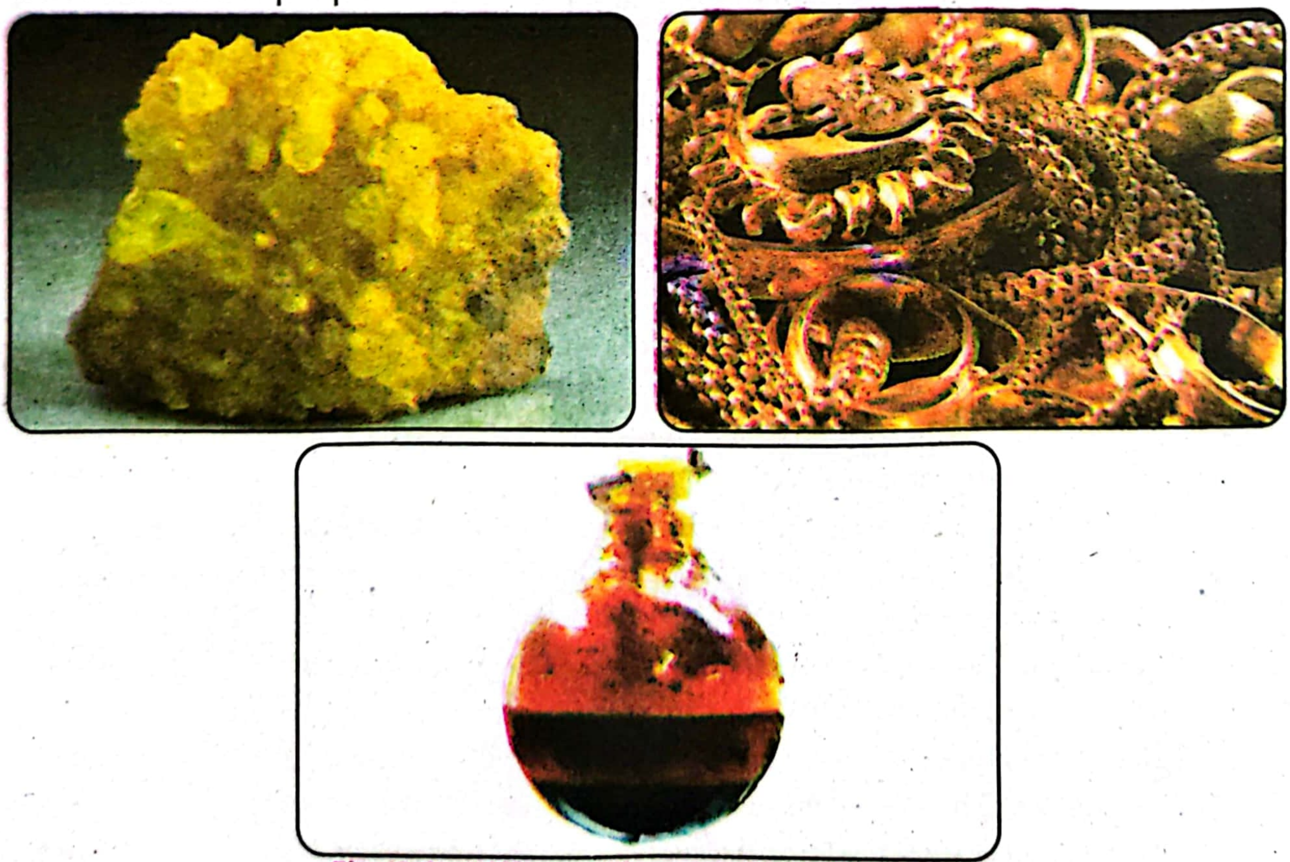


Fig (2.1) Different Elements look different

Write down the names of the elements shown in the above pictures.

2.1 Structure of Atom

The idea of atom was first proposed in Greece when the philosopher Democritus declared that all matter is made of tiny particles. He named this particle as atom, a particle that cannot be further subdivided. Around 1800, the English chemist John Dalton did many experiments with compounds that provided the evidence for the existance of atoms. In the last decade of nineteenth century, a group of scientists were trying to pass electricity through gases at reduced pressure. During the course of these experiments known as 'Discharge Tube Experiments', they discovered that atoms are no longer the smallest particles of matter; rather there exist particles that are even smaller than atoms. In other words, atoms are composed of negatively charged particles called electrons and positively charged particles called protons. It was also discovered that a proton is 1836 times heavier than an electron. In a discharge tube, the presence of the negatively charged particles was ascertained because of their deflection towards the positive plate in an electric field. Similarly, the presence of positively charged particles was confirmed due to their deflection towards the negative plate.

Discovery of Electrons

A discharge tube is a hard glass tube provided with two metallic electrodes and a vacuum pump to evacuate the gas present in it Fig (2.2). When a very high voltage is applied to a gas at a very low pressure present in a glass tube the glass surface behind the positive electrode started to glow, due to the rays emitted from the cathode.

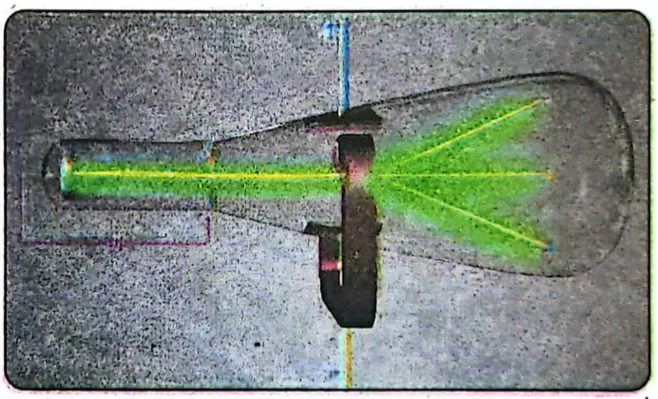


Fig (2.2) Discharge Tube Experiment Cathode Rays

These rays were named as cathode rays. In 1897, British physicist Joseph John Thomson studied the properties of cathode rays by passing them through the oppositely charged electric plates. It was observed that cathode rays bent towards the positively charged plate showing that they carry negative charge. Thomson also installed two magnets on either side of the discharge tube and noticed that cathode rays were also diverted by the magnetic field. Thomson used the findings of his experiments to calculate the mass to charge ratio of cathode rays which finally proved that cathode rays are in fact, negatively charged material particles. These particles were later named as electrons. It was also shown that electrons are the subatomic particles of all elements.

Discovery of Proton

The presence of positively charged particles in an atom had been first observed by E.Goldstein in 1886. It was based on the concept that atoms are electrically neutral having same number of positive and negative charges. He performed a series of experiments with a gas-discharge tube having a perforated cathode. A new type of rays were produced from the anode which moved towards the cathode. He called these new rays as canal rays or anode rays.

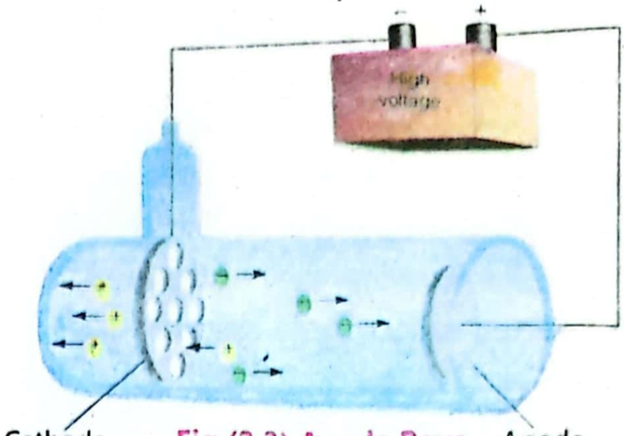


Fig (2.3) Anode Rays

The properties of these rays seemed to vary depending on the gas used in the discharge tube. In fact what he discovered was gas ions and this also included hydrogen ions (H^+). Goldstein at that time knew nothing about its significance.

In 1917, Rutherford performed experiments which proved that the hydrogen nucleus is present in other nuclei. Rutherford thought that a hydrogen nucleus or a proton must be the fundamental building block of all nuclei and also possibly a new fundamental particle as well.

Discovery of Neutron

Later on, in 1933, another particle neutron was also discovered, which is known to carry no charge. The mass of a neutron is almost the same as that of a proton.

These three particles i.e. electron, proton and neutron were given the name fundamental particles and are shown to be present in all atoms (except ordinary hydrogen which does not have a neutron) irrespective of the fact that these atoms behave very different from one another. It was, however, also shown that the number of these particles is different in different atoms.

After the discovery of these particles, a very important question arose as to how these particles are arranged in a tiny place called atom. In other words, what is the structure of an atom?

Discovery of Nucleus

Lord Rutherford, in 1911, provided answer to this question. He carried out a remarkable experiment in which he hit a stream of special type of particles to a

Interesting Information!

Although the nucleus is less than one hundred-Thousandth (1/100,000) of the size of the atom, it contains more than 99.9% of the mass of the atom.

very thin gold foil. From this experiment he concluded that an atom has two portions. A tiny central portion which he called as nucleus and a relatively large area surrounding this, which he called extra nuclear portion.

It was also discovered that almost all the mass of an atom is concentrated in the nucleus because both the heavy particles i.e. protons and neutrons are found to be present here. In the nucleus these two particles are held together by a strong nuclear force.

Particle	Charge	Mass
Electron	-1.6022×10^{-19} C	9.109×10^{-31} Kg
Proton	$+1.6022 \times 10^{-19}$C	1.673×10^{-27} Kg
Neutron	0.0	1.675×10^{-27} Kg

Bohr's Atomic Model

In 1913, Niels Bohr proposed a model for the hydrogen atom which is called Bohr's atomic model. According to this model the electron can revolve around the nucleus of the atom in specific paths called orbits or shells. When the electron is revolving in one of these orbits, its energy is fixed. Also, when the electron is present in the orbit which is closest to the nucleus, its energy is minimum and it is called the ground state of the atom. The orbits that are further away from the nucleus possess successively greater energy. The electron is not allowed to occupy a space in between the orbits.

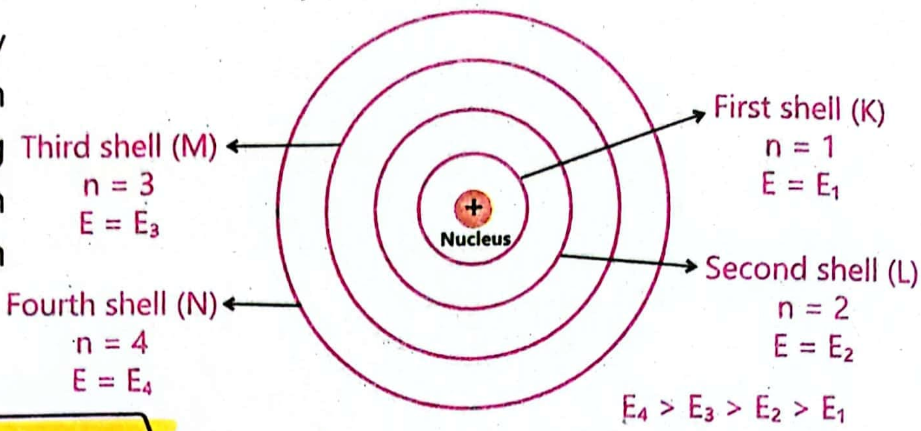
Since electron present in each shell has a fixed energy, these shells are also named as energy levels. The shell which is nearest to the nucleus is called first shell or K shell and the electron present in it has a fixed value of energy. Similarly, the second shell will also be at a definite distance from the nucleus which will, of course, be greater than the first shell and electron present in it will also possess greater value of energy. Similarly, electron may also be present in the third or higher shells.

Each shell is further sub-divided into sub-shells or orbitals. The number of sub-shells present in a shell is equal to the value of n for that shell. For the first shell, (n = 1), it will, therefore, have only one sub-shell which is called **s** sub-shell. For (n = 2), there will be two sub-shells **s** and **p**. The second shell will, therefore, have two sets of sub-shells. The third shell (n = 3) has three sub-shells **s, p** and **d**. The fourth shell consists of four sub-shells **s, p, d** and **f**.

Interesting Information!

The size of an atom is so small that it is not possible to see it with naked eye. However, a transition electron microscope can be used to see atoms.

According to modern approach, electrons are like charged clouds whose location around the nucleus cannot be predicted with one hundred percent certainty. All we can say is that there is a certain probability of finding the electron at a certain probable distance from the nucleus.



Interesting Information!

The largest atom cesium is approximately nine times bigger than the smallest atom helium.

In order to find out the number of electrons which can be accommodated in these extra nuclear shells, the scientists have devised a formula called ($2n^2$) formula where n can have values 1,2, 3 ... and so on and they represent the number of shells. Shells have also been named as **K**, **L**, **M**, **N** .... and so on. When the value of n is one, it means first or K shell and the maximum number of electrons which can be present in this shell is ($2 \times 1^2 = 2$). K shell can, however, have less than two electrons or it may not have any electron.

Similarly when (n = 2), it means second or L shell and the maximum number of electrons it can accommodate is ($2 \times 2^2 = 8$). For (n = 3) or M shell, it can accommodate 18 electrons at the most. This process will go on until the electrons present in an atom are finished.

The maximum number of electrons which can be accommodate in sub-shells **s**, **p**, **d** and **f** are 2, 6 10 and 14 respectively. In the first shell there are 2 electrons which shall go to (**s**) sub-shell. In the second shell, 8 electrons will be further sub-divided, s-subshell will have 2 electrons and 6 electrons will be accommodated in p-subshell.

2.2 Atomic Number and Mass Number

Electrons, protons and neutrons are called the fundamental particles of all types of matter. In other words, the atoms of all the elements present in this world contain same electrons, protons and neutrons. However, an atom of one element differs from an atom of another element because it contains different number of the fundamental particles.

The number of protons present in the atoms of an element is always fixed

and it is called the atomic number of that element. Since an atom, as a whole, is electrically neutral, the number of electrons present in an atom will be the same as the number of protons. Atomic number of an element is represented by **Z**. Atomic number of an element is unique to that element and the element is identified by this number. In a periodic table of elements, the elements are arranged according to ascending order of their atomic numbers.

The total number of protons and neutrons present in an atom almost accounts for the total mass of that atom and hence it is called its nucleon number or mass number. It is represented by **A**. The mass of electron being very small is not included in the mass number.

Just like atomic number, the atom of an element may also be identified from its mass number. For example, the number of protons present in an oxygen atom is 8, so its atomic number is 8 while the total number of protons and neutrons present in it is 16, so its mass number is 16. Information about the atomic number and the mass number is often included with the symbol of any element. The atomic number is written as a left subscript while the mass number as a left superscript. So, oxygen atom would be symbolized as $^{16}_{8}O$. Similarly, carbon atom symbolized as $^{12}_{6}C$ will have 6 protons and 6 neutrons.

The number of neutrons **N** present in an atom can be calculated if its atomic number **Z** and mass number (A) are known.

$$N = A - Z$$

Thus, the number of neutrons in sulphur atom, symbolized as $^{32}_{16}S$, can be calculated as $32 - 16 = 16$.

Interesting Information!

Changing the number of electrons of an element forms ions, while changing the number of neutrons of an element forms isotopes.

Sample Example

Calculate the number of neutrons, protons and electrons in barium $^{137}_{56}Ba$

Solution: Each barium atom will have 56 protons and 56 electrons. The number of neutrons in barium will be calculated as follows:

$$N = A - Z$$

$$N = 137 - 56 = 81$$

So each $^{137}_{56}Ba$ atom will have 81 neutrons, 56 protons and 56 electrons.

Exercise

Calculate the number of neutrons, protons and electrons in the following atoms.

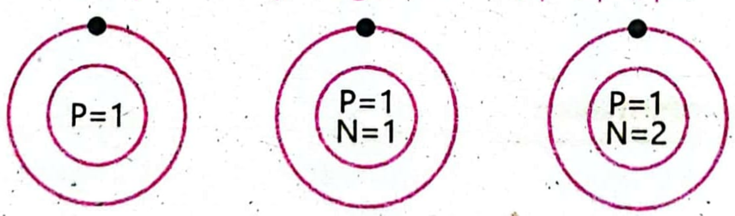
$^{195}_{78}Pt$, $^{55}_{25}Mn$, $^{127}_{53}I$

2.3 Isotopes and their Masses

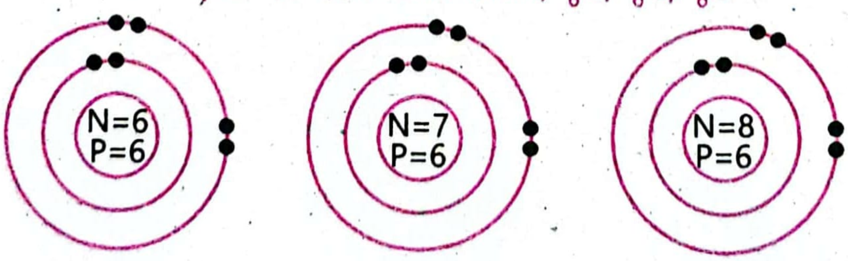
All the atoms of an element must necessarily have the same atomic number, but their mass number may vary depending upon the number of neutrons present in the nucleus. Atoms of the same element having different number of neutrons in their nuclei are called isotopes. For example, element carbon has three isotopes as its atoms have six, seven and eight neutrons in their nuclei. These isotopes are represented as $^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$. Similarly, hydrogen exists as three isotopes, Hydrogen, Deuterium and Tritium represented by $^{1}_{1}H$, $^{2}_{1}H$ and $^{3}_{1}H$. Hydrogen ($^{1}_{1}H$) is the only atom which does not have a neutron. Since the chemical properties of the elements are determined by the number of electrons, all three isotopes will show almost the same chemical behaviour, although their physical properties may be different.

$^{2}_{1}H$ has twice the mass of $^{1}_{1}H$ while the mass of $^{3}_{1}H$ is thrice as the mass of $^{1}_{1}H$. Similarly, the masses of three different isotopes of carbon are different.

Isotopes of Hydrogen Atom, $^{1}_{1}H$, $^{2}_{1}H$, $^{3}_{1}H$



Isotopes of Carbon Atom, $^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$



Exercise: Why isotopes of an element show same chemical properties while their physical properties are different?

Radioactive Isotopes

The isotopes of the same element do not have the same physical properties. Several isotopes of the same elements exist whose nuclei are unstable. They emit excess energy in the form of radiation. This process of emission of radiation is called radioactivity and the isotope which emits energy is called radioactive isotope. Tritium $^{3}_{1}H$ is a radioactive isotope and the other two isotopes of hydrogen are stable and do not emit any radiation.

When a radioactive element emits radiation, it is transformed into another chemical element. This process is called radioactive decay. This new element may be stable or may be radioactive so that it also emits radiation.

$$^{238}_{92}U \text{ (Uranium)} \longrightarrow {}^{234}_{90}Th \text{ (Thorium)} + \text{energy}$$

$^{234}_{90}$Thorium is unstable and further disintegrates to give $^{231}_{91}$Pa (Protactinium)

$$^{234}_{90}Th \text{ (Thorium)} \longrightarrow {}^{231}_{91}Pa \text{ (Protactinium)} + \text{energy}$$

$$^{210}_{83}Bi \text{ (Bismuth)} \longrightarrow {}^{206}_{81}Tl \text{ (Thallium)} + \text{energy}$$

Interesting Information!

Every year, our body replaces about 98% of its atoms.

Applications of Radioactive Isotopes

Radioactive isotopes are useful in medical imaging. Doctors use them to diagnose the disease by injecting the patient with a small amount of radioactive fluid. Technetium – 99m is used for diagnostic imaging across human organs like brain, lungs, etc. Doctors use a special camera to watch how the radioactive fluid moves.

Exercise

1. Why does a radioactive isotope emit radiation?
2. Give an example of a radioactive isotope which disintegrates to give a stable atom.

Interesting Information!

Gallium has many interesting properties. Its melting point is below body temperature.so it is liquid at room temperature. It does not evaporate.

Radiocarbon dating is a method for finding out the age of an historical object containing organic material with the help of radioactive isotope of carbon $^{14}_{6}C$. The method involves measuring the proportion of ^{14}C in a sample from a dead plant or animal like a piece of wood or a bone which provides information that can be used to calculate when an animal or plant died. The older the sample is, the less ^{14}C is to be detected.

Radioactive isotopes are used to test the strength of metals and concrete mixture. They are used to generate cheap nuclear power and to find oil fields. In medicine they are used to diagnose and treat many medical conditions and diseases, including cancer and thyroid disorders.

Ionization of Atoms by a Radioactive Source

Radiation emitted from a radioactive source causes atoms to ionize. For example, radiation emitted by a radioactive element radium -226 can remove electron or electrons from the atom. However, this ionizing radiation should have enough energy to remove the tightly bound electron from the orbit of an atom. Electron can be lost because an ionizing radiation collides with the atom and forces the electron away from the atom. It an atom of sodium is hit by an ionizing radiation, it may lose an electron. This process converts the atom into a positively charged ion (cation).

$$Na \xrightarrow{\text{Energy}} \underset{\text{Cation}}{Na^{+}} + e^{-}$$

Interesting Information!

The same elements occur everywhere in the universe. Matter on Mars or in any other planet consists of the same elements found on Earth.

2.4 Relative Atomic Mass

Ever since the existence of atom was recognized, the chemists were trying to find out a method which would allow them to compare the masses of different atoms. It was necessary because without knowing the relative masses of atoms, we would not know in which ratio of masses we mix the reactants to carry out a chemical reaction. In 1961, the chemists adopted a new scale for the

measurement of the relative masses of atoms. The unit mass on this scale is 1/12th of the mass of lighter isotope of carbon taken as 12. The mass of one atom of carbon on this scale is exactly 12 and the masses of the other atoms are measured relative to this unit. The relative atomic mass of an element is thus defined as the mass of an atom of that element relative to the mass of light isotope of carbon taken as 12. The relative atomic masses of elements are expressed in atomic mass unit (amu). It is defined as one-twelfth the mass of an atom of carbon-12.

$$1\text{amu} = 1.67377 \times 10^{-27}\ \text{kg}$$

For example, the mass of one atom of hydrogen-1 on atomic mass scale is 1.007 amu, similarly the mass of one atom of sulphur-32 is 31.972 amu.

Exercise

How would you compare the masses of the atoms of Mg and Cl?

Calculation of Relative Atomic Mass from Isotopic Abundance

An element usually consists of a few different isotopes with different mass numbers. These mass numbers are called relative isotopic masses. Each isotope will also have its own naturally occurring abundance which is called isotopic abundance.

Relative atomic mass of an element can be calculated from the relative isotopic masses (m) and isotopic abundances (p) by the following formula

$$\text{Relative atomic mass} = \frac{m_1p_1 + m_2p_2 + m_3p_3\ldots\ldots\ldots\ldots}{100}$$

The element Krypton (Kr) has five isotopes. Their relative isotopic masses and isotopic abundances are shown in the following Table (2.2).

Table (2.2) Isotopic Abundances of Krypton

Relative Isotopic Mass	Isotopic Abundance
80	2.0%
82	12.0%
83	12.0%

84	57.0%
85	17.0%

Sample Example 1:

Relative atomic mass of krypton

$$= \frac{80 \times 2.0 + 82 \times 12.0 + 83 \times 12.0 + 84 \times 57.0 + 85 \times 57.0}{100}$$

$$= 83.7$$

Sample Example 2:

Calculate the relative atomic mass of light isotope of chlorine when its relative atomic mass is taken as 35.45.

Relative isotopic abundance of Cl-37 = 24.23%

Relative isotopic abundance of light isotope of chlorine = 75.77%

Relative atomic mass of chlorine = 35.45

Relative atomic mass of chlorine $= \frac{Cl \times 75.77 + 37 \times 24.23}{100}$

$$35.45 = \frac{Cl \times 75.77 + 37 \times 24.23}{100}$$

$$3545 = Cl \times 75.77 + 37 \times 24.23$$

$$3545 - 896.51 = Cl \times 75.77$$

$$2648.49 / 75.77 = Cl$$

$$34.95 = Cl$$

Relative atomic mass of light isotope of chlorine is 34.95.

Exercise

Calculate the relative atomic mass of Lead (Pb). Isotopic abundances of isotopes are 2.0, 24.0, 22.0, 52.0 respectively.

^{204}Pb ^{206}Pb ^{207}Pb ^{208}Pb

Key Points

1. An English chemist, John Dalton provided the evidence for the existence of atoms.
2. Discharge tube experiments showed that atoms are no longer smallest particles of matter. Rather they are made up of still smaller particles called electron and proton. Neutron was discovered separately.
3. Electrons, protons and neutrons are shown to be present in all the elements irrespective of the fact that the elements behave very differently. Different elements, however, contain different number of these particles.
4. Lord Rutherford discovered that all atoms have a central part which he named as nucleus. The protons and neutrons are present in this nucleus while the electrons are revolving around the nucleus.
5. An atom being electrically neutral contains the same number of electrons and protons.
6. According to Bohr's atomic model, the electron revolves around the nucleus in definite circular paths called orbits or shells. As long as electron revolves in an orbit its energy remains fixed. An electron will have more energy if it is located away from the nucleus.
7. The number of protons present in the nucleus of an element is called the atomic number of that element.
8. The total number of protons and neutrons present in the nucleus of an element in called its mass number.
9. Isotopes are the atoms of the same element which have the same number of protons but different number of neutrons.
10. Isotopes of an element have same chemical properties but they differ in their physical properties.
11. Isotopes of an element may be stable or radioactive. Radioactive isotopes have many useful applications in medicine.
12. Radioactive isotopes have unstable nuclei and they throw out radiation.
13. Relative atomic mass of an element can be calculated from the relative isotopic masses of this element and their isotopic abundances.

Exercise

1. Tick (✓) the correct answer.

(i) How many electrons can be accommodated at the most in the third shell of the elements?

(a) 8 (b) 18 (c) 10 (d) 32

(ii) What information was obtained from discharge tube experiments?

(a) Structure of atom was discovered.
(b) Neutrons and protons were discovered.
(c) Electrons and protons were discovered.
(d) Presence of nucleus in an atom was discovered.

(iii) Why have isotopes not been shown in the periodic table?

(a) Periodic table cannot accommodate a large number of isotopes of different elements.

(b) Some of the isotopes are unstable and they give rise to different elements

(c) All the isotopes have same atomic number; so there is no need to give them separate places.

(d) Isotopes do not show periodic behavior.

(iv) Which particle is present in different number in the isotopes?

(a) Electron (b) Neutron

(c) Proton (d) Both neutron and electron

(v) In which isotope of oxygen there are the equal number of protons, electrons and neutrons?

(a) ^{17}O (b) ^{16}O (c) ^{18}O (d) Non of these

(vi) What will be the relative atomic mass of nitrogen given the abundances of its two isotopes, ^{14}N and ^{15}N are 99.64 and 0.35 respectively.

(a) 14.0210 (b) 14.0021 (c) 14.2100 (d) 14.1200

(vii) How is radiocarbon dating useful for archeologists?

(a) It helps determine the age of organic matter.

(b) It helps determine the composition of matter.

(c) It helps determine the usefulness of matter.

(d) It helps determine whether the matter is radioactive or not.

(viii) What does keep the particles present in the nucleus intact?

(a) Particles are held together by strong nuclear force.

(b) Particles are held together by weak nuclear force.

(c) Particles are held together by electrostatic force.

(d) Particles are held together by dipolar force.

(ix) How do electrons keep themselves away from the oppositely charged nucleus?

(a) By keeping themselves stationary

(b) By revolving around the nucleus

(c) Due to their wave-like nature

(d) A magnetic field around the nucleus keeps them away

(x) Rubidium consists of two isotopes ^{85}Rb and ^{87}Rb. The percent abundance of the light isotope is 72.2% What is the percent abundance of the heavier isotope? Its atomic mass is 85.47

(a) 15% (b) 27.8% (c) 37% (d) 72%

2. Questions for Short Answers

i. Why is it said that almost all the mass of an atom is concentrated in its nucleus?

ii. Why are elements different from one another?

iii. How many neutrons are present in $^{210}_{83}Bi$?

iv. Why is tritium ($^{3}_{1}H$) a radioactive element?

v. How can an atom absorb and envolve energy?

3. Constructed Response Questions

i. Why does the energy of electron increase as we move from first shell to second shell?

ii. Why is it needed to lower the pressure of the gas inside the discharge tube?

iii. What is the classical concept of an electron? How has this concept changed with time?

iv. Why the nuclei of the radioactive elements are unstable?

v. During discharge tube experiments, how did the scientists conclude that the same type of electrons and protons are present in all the elements?

4. Descriptive Questions

i. Explain the structure of a hydrogen atom.

ii. How does the theory of atomic structure explain the ionization of atoms by a radioactive isotope?

iii. What is radioactivity? Explain any three applications of radioactive isotopes.

iv. Find out the relative atomic mass of mercury from the following data.

Isotope		Relative Abundance	Isotope		Relative Abundance
^{196}Hg	=	0.0146%	^{199}Hg	=	16.34%
^{198}Hg	=	10.02%	^{200}Hg	=	23.13%
^{201}Hg	=	13.22%	^{202}Hg	=	29.80%
^{204}Hg	=	6.85%			

5. Investigative Questions

i. How can scientists synthesize elements in the laboaratory?

ii. A system just like our solar system exists in an atom. Comment on this statement.

Chapter 3

Chemical Bonding

Student Learning Outcomes

After studying this chapter, students will be able to:

- Describe that noble gas electronic configuration, octet and duplet rules help predict chemical properties of main group elements
- Compare between the formation of cations and anions
- Account for the electropositive and electronegative nature of metals and non-metals.
- Define ionic, covalent, coordinate covalent and metallic bonds
- Differentiate between ionic compounds and covalent compounds. (The following points need to be included in the respective definitions:

  a. Ionic Bond as strong electrostatic attraction between oppositely charged ions

  b. Covalent bond as strong electrostatic attraction between shared electrons and two nuclei

  c. Metallic bond as strong electrostatic attraction between cloud/sea of delocalized electrons and positively charged cations)
- Explain the properties of compounds in terms of bonding and structure
- Compare uses and properties of materials such as strength and conductivity as determined by the type of chemical bond present between their atoms.
- Interpret the strength of forces of attraction and their impact on melting and boiling points of ionic and covalent compounds.
- Justify the availability of free charged particles (electrons or ions) for conduction of electricity in ionic compounds (solid and molten) covalent compounds and metallic bonds.
- Recognize that some substances can ionize when dissolved in water. (e.g. acids dissolves in water and conduct electricity)
- Justify the suitability of usage of graphite, diamond and metals for industrial purposes (Some examples may include: **a.** graphite as lubricant or an electrode **b.** diamond in cutting tools **c.** metals for wires, and sheets)
- Draw the structure of ionic and covalent compounds along with their format ion. (some examples can include: **a.** ionic bonds in binary compounds such as NaBr, NaF, $CaCl_2$ using dot-and- cross diagrams and Lewisdot structures simple molecules including H_2, Cl_2, O_2, N_2, H_2O, CH_4, NH_3, HCl, CH_3O H, C_2H_4, CO_2, HCN, and similar molecules using dot and-cross diagrams and Lewis- dot structures)

3.1 Why do atoms form chemical bonds?

Atoms have a tendency to decrease their energy. They can do this by combining with other atoms. It is a natural phenomenon because it increases the stability of atoms.

How do atoms succeed in lowering their energy? The early chemists had started thinking about this a long time ago. They finally succeeded to get an answer only when the noble gases He, Ne, Ar, Kr and Xe were discovered. Helium has two electrons in its outermost shell while all other noble gases have eight electrons in their outermost shells. We also know about these gases that neither their atoms combine with themselves nor with other atoms. The probable reason for this lack of reactivity was their stability. It was suggested that these gases were stable due to the presence of two electrons in helium and eight electrons in the outermost shells of the rest of the gases. This gave rise to a principle that having two electrons (for hydrogen and helium which have only the first shell) or eight electrons in the outermost shell meant stability and hence unreactivity as well. This principle was named as Duplet or Octet Rule.

The discovery of duplet or octet rule was followed by another similar suggestion that atoms form bonds because they would like to lower their energy by completing their duplet or octet. For example, for sodium atom it is easy to lose one electron and stabilize itself than to gain seven electrons while completing its octet. Sodium atom, therefore, adopts the energetically easier path and loses its electron to form a bond. In the same way, it is energetically favourable for hydrogen atom to lose one electron to become proton (H^+) or gain one electron to become hydride ion (H^-). In the latter case, it completes its duplet.

Alkali and alkaline earth metals are therefore expected to be electropositive metals which will form bonds with electronegative elements like oxygen and chlorine. Although, in the beginning, octet rule played a significant role in understanding the nature of a chemical bond, yet further investigations found it to be less important.

3.2 Chemical Bond

A chemical bond is a force of attraction between atoms which holds them together in the form of a molecule or a compound.

When atoms of different substances approach each other, there are two possibilities. They may attract or repel each other. If the forces of attraction between them dominate the forces of repulsion, the energy of the system gets

lowered and as a result the two atoms will react to form a new molecule. Conversely, the two atoms simply move away from each other.

Important Information!

The arrangement of electrons around the nucleus of an atom in shells and sub-shells is called electronic configuration.

Types of Bonds

We shall consider here three types of bonds.

(1) Ionic bond
(2) Covalent bond
(3) Coordinate covalent bond

3.2.1 Ionic Bond

This chemical bond is formed as a result of the tendency of atoms to lose or gain electron or electrons to acquire the electronic configuration of the nearest noble gas because this is a more stable electronic structure. Let us take the example of the formation of a simple and important compound, sodium chloride. This compound is formed when the elements sodium and chlorine react chemically. The electronic configurations of these elements are shown in Fig (3.1).

	1st shell	2nd shell	3rd shell
$_{11}Na$	2	8	1
$_{17}Cl$	2	8	7

Fig (3.1): Electronic Configurations of Sodium and Chlorine

An electron from the outermost shell of sodium atom is transferred to the outermost shell of chlorine atom and in doing so, both these atoms acquire the electronic configurations of their nearest noble gases. (Fig 3.2)

$$Na \longrightarrow Na^+ + e^-$$
$$e^- + Cl \longrightarrow Cl^-$$
$$Na + Cl \longrightarrow Na^+Cl^-$$

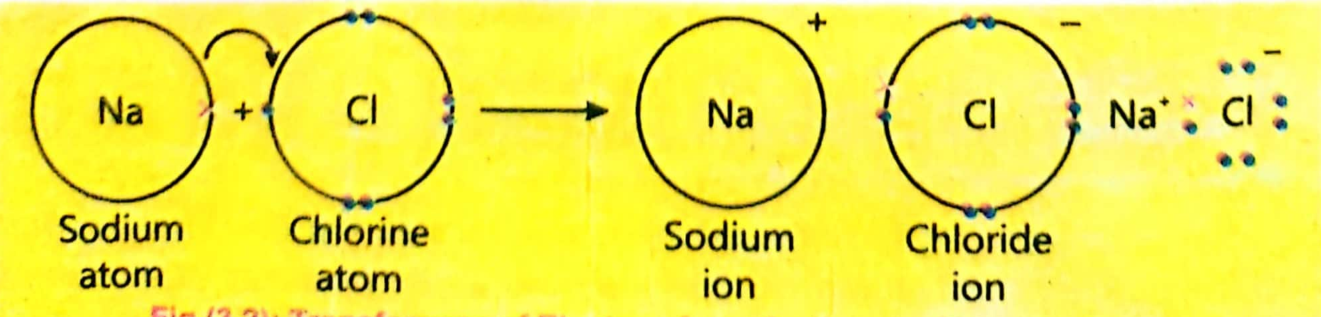


Fig (3.2): Transference of Electron from Sodium to Chlorine Atom

Similarly, sodium also reacts with fluorine and bromine to give sodium fluoride and sodium bromide respectively.

It should be noted here that an electron or electrons, which take part in a chemical reaction, come only from the outermost shells of the atoms. Sodium chloride, formed as a result of the chemical reaction mentioned on the previous page contains the positively charged sodium ions (Na^+) and the negatively charged chloride ions (Cl^-). These oppositely charged ions are then held together by the electrostatic force of attraction. The chemical bond, thus formed, is called an Ionic or an Electrovalent Bond and the compounds having such a bond are called ionic compounds.

Calcium, an alkaline earth metal, loses two electrons to form calcium chloride ($CaCl_2$). Fig (3.3)

$$Ca \longrightarrow Ca^{2+} + 2\bar{e}$$

$$2Cl + 2\bar{e} \longrightarrow 2Cl^-$$

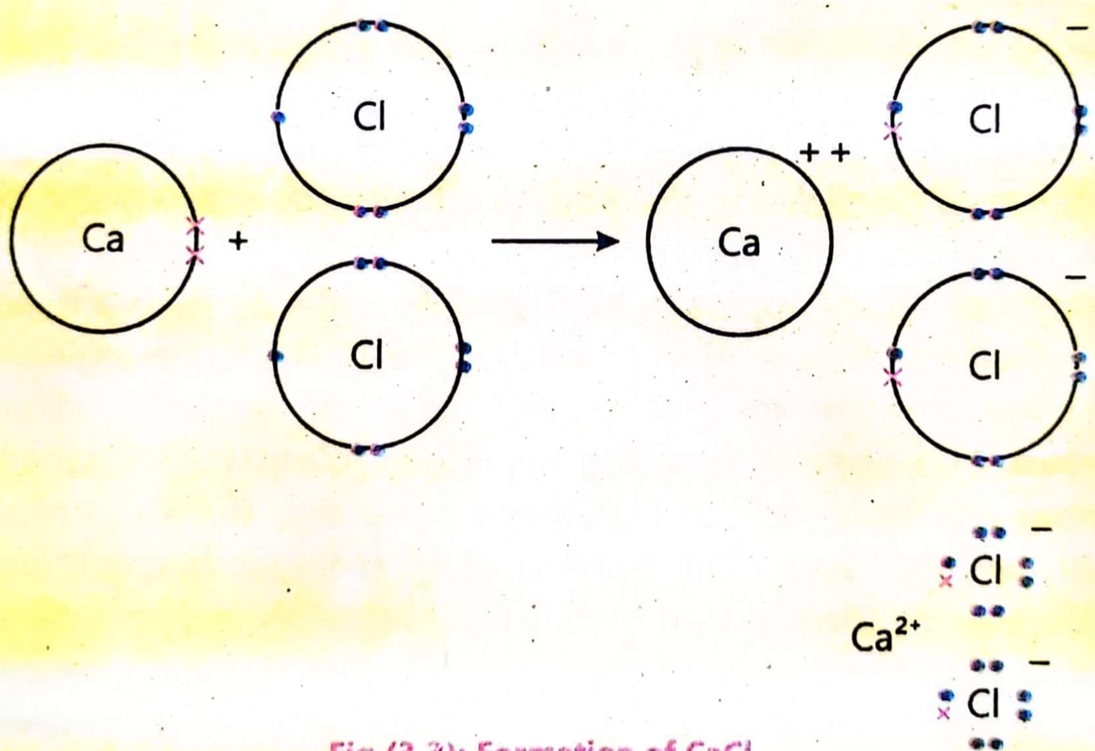


Fig (3.3): Formation of $CaCl_2$

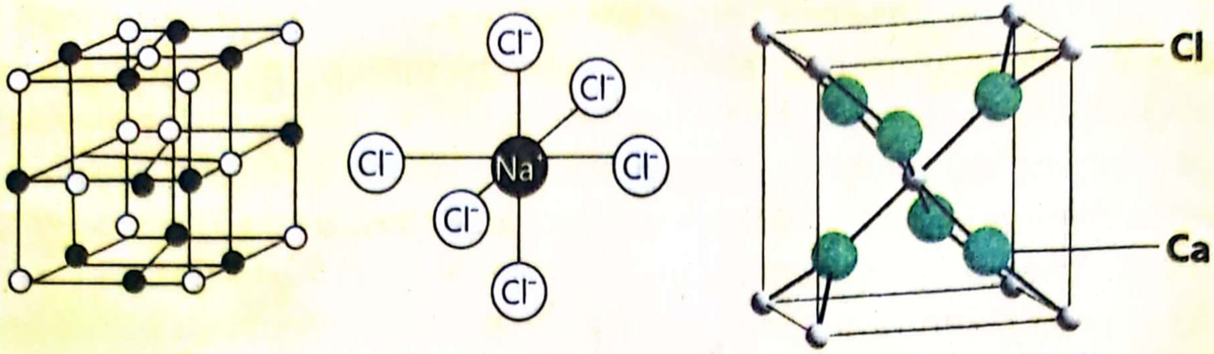
These ions then surround each other three dimensionally to form a crystal lattice.

Examples of ionic compounds are KCl, Mg F, NaF, KBr, CaF_2

Exercise

1. What types of elements form ionic bonds?
2. What are the conditions for an ionic bond to be formed?

Figures of crystal lattices of NaCl and $CaCl_2$. (Fig 3.4)



Crystal Lattice of NaCl

Crystal lattice of $CaCl_2$

Fig (3.4) Crystal Lattices of NaCl and $CaCl_2$

An ionic bond is therefore a bond which is formed by the complete transference of electron or electrons from one atom to another atom.

3.2.2 Covalent Bond

During the formation of an ionic bond, the atoms lower their energy by the transference of an electron (s) and thus acquire the electronic configurations of the nearest noble gas. However, it is not the only way by which atoms can lower their energy. Some atoms decrease their energy by mutually sharing their electrons. This can be explained as follows:

When two atoms approach each other in order to form a bond, they undergo important changes in their energy. The electrons belonging to one atom will come under the attractive influence of the nucleus of the other atom. This is the new force of attraction and will be responsible for lowering the energy. The electrons and the nucleus of one atom will also repel the electrons and the nucleus of the other atom. This is the force of repulsion and will obviously increase the energy. The two atoms will bring themselves at such a distance so that the attractive forces dominate the repulsive forces. The total energy at this distance will be minimum and thus a stable molecule is formed. **A covalent bond is therefore a bond formed by the mutual sharing of an electron pair provided by the bonded atoms. This is called a single covalent bond.**

In some compounds, the atoms share two electrons each to form a double covalent bond. In the same way atoms can share three electrons each to form a triple covalent bond. Double and triple covalent bonds have two and three electron pairs respectively which are mutually shared between the two atoms. A single covalent bond is represented by a single line(–), a double covalent bond is represented by two lines (=) while a triple

covalent bond is represented by three lines(≡). The mutually shared electrons may be shown by a dot or a cross. The formation of single, double and triple covalent bonds in different molecules is explained in the examples shown in Fig (3.5).

Exercise

1. What type of elements form covalent bond?
2. How covalent bond is different from an ionic bond?

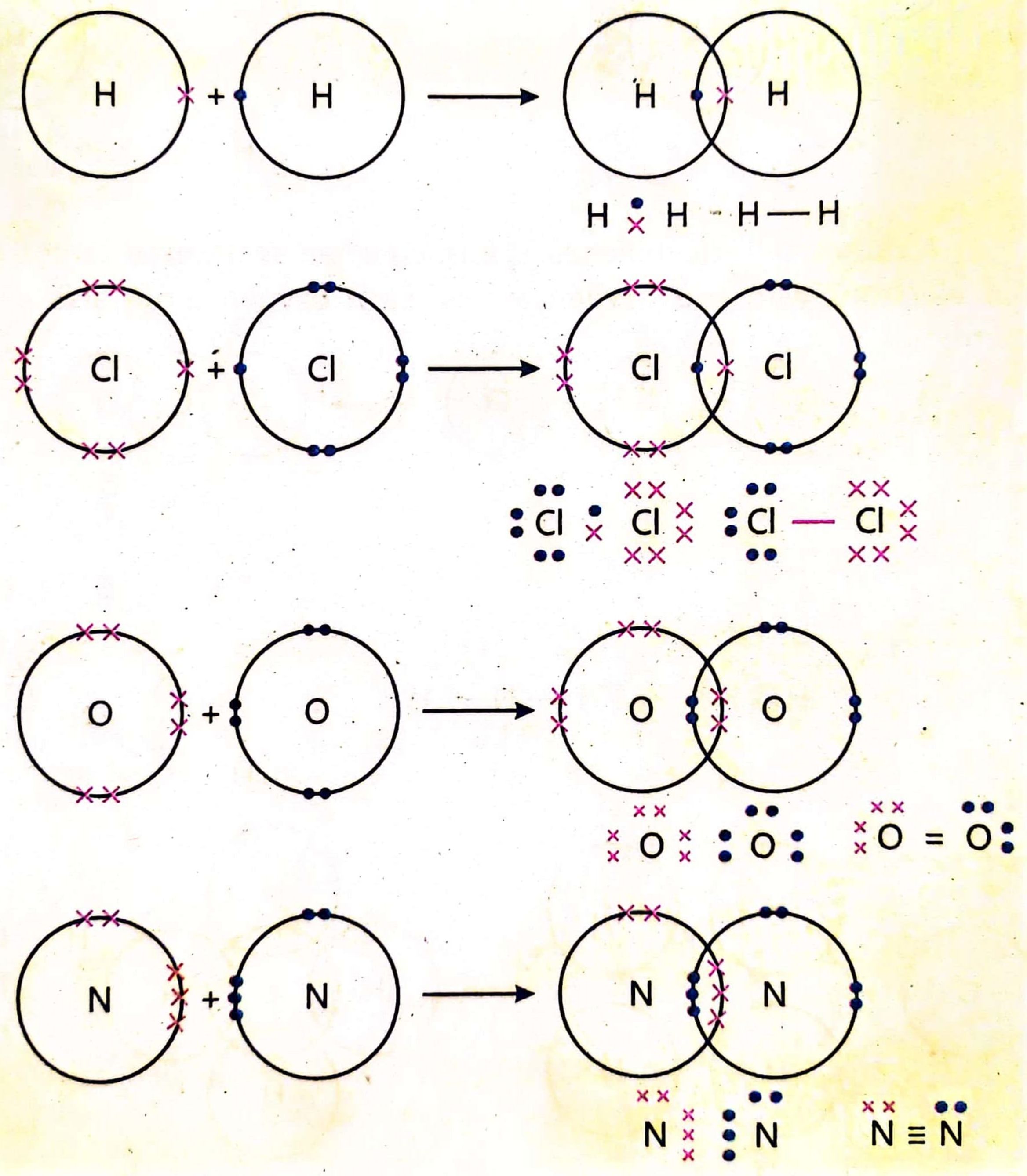
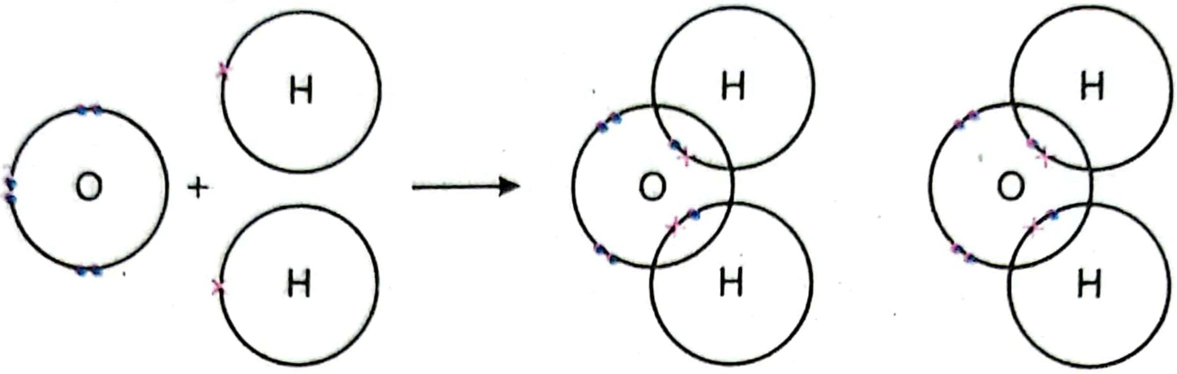


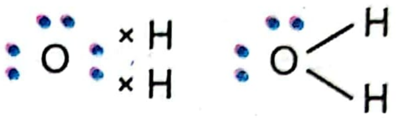
Fig (3.5): Formation of Single, Double and Triple Covalent Bonds

Formation of Covalent Compounds

Water

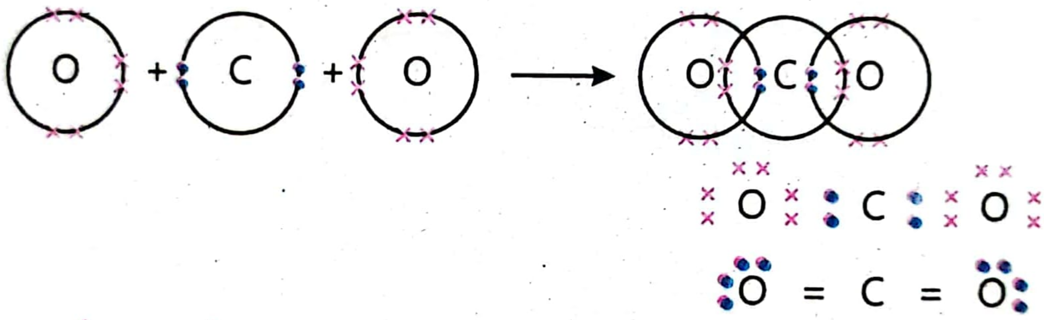
A water molecule is formed when two hydrogen atoms share their electrons separately with the electrons of one oxygen atom.



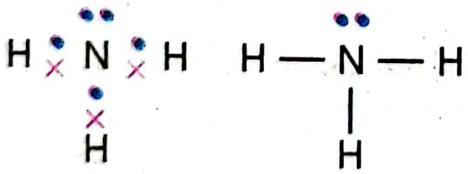


Carbon dioxide

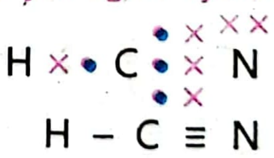
A carbon dioxide molecule is formed when an atom of carbon shares its four electrons with two oxygen atoms. Each oxygen atom also shares two electrons.



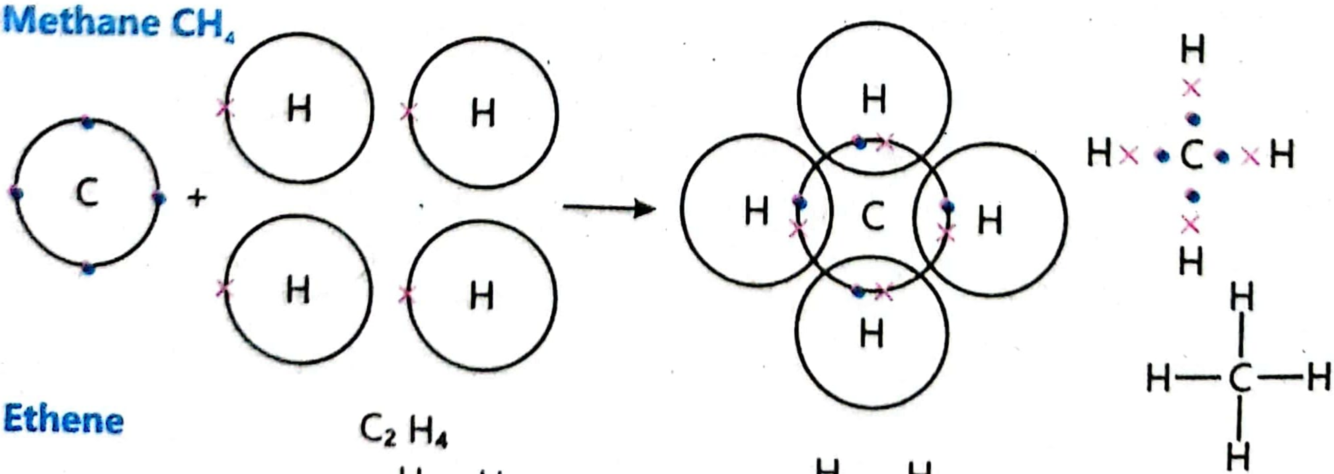
Ammonia



Hydrogen cyanide

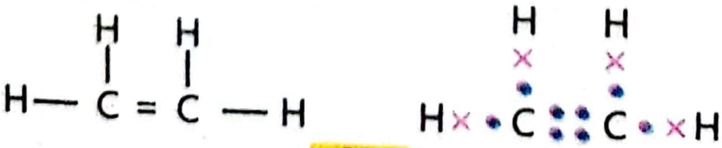


Methane CH_4



Ethene

C_2H_4



Methanol

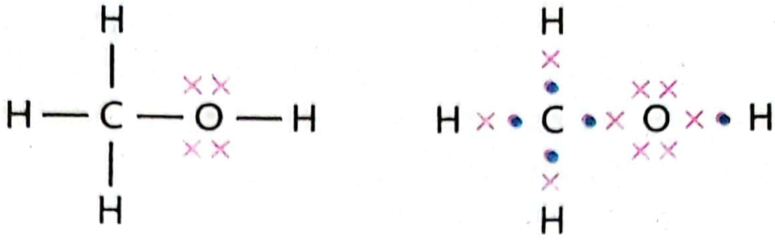


Fig (3.6): Formation of Covalent Compounds

Exercise

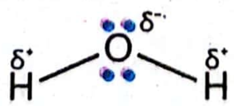
Draw electron dot and cross structure of the following compounds.
SiH_4, PCl_3, SO_2, SO_3

It is quite clear from the examples shown above that after mutually sharing their electrons, the bonded atoms acquire the electronic configurations of the nearest noble gas.

Polar and Non-polar Covalent Bonds

When a covalent bond is formed between two identical atoms, the shared pair of electrons lies at the middle of the two bonded atoms. This gives rise to a pure covalent bond which is essentially non-polar in character. For example, covalent bonds in hydrogen and chlorine molecules are non-polar in nature. Contrary to this when a covalent bond is formed between two non-identical atoms, the shared pair of electrons bend towards the atom which has more electronegativity. As a result a partial nagative charge is created on the more electronegative atom and an equal partial positive charge is created on the less electronegative atom. A polar covalent bond is thus formed. For example, in HCl molecule, the shared pair of electron bends toward the more electronegative chlorine atom creating a partial negative charge on it. The covalent bond present in HCl molecule is thus a polar covalent bond.

Similarly the covalent bonds present in water are also polar covalent bonds.



3.2.3 Coordinate Covalent Bond

Coordinate covalent bond is a type of covalent bond in which the shared electron pair is donated by one atom only. This bond is formed when a species has an electron pair to donate to another species. The species which donates the electron pair, is called a donor while that which accepts it is called an acceptor. An arrow head (→) pointing towards the acceptor represents this type of bond. Following examples will help to explain this bond.

Hydroxonium Ion (H_3O^+)

Acids provide protons (H^+) when dissolved in water. This proton has an empty outer shell and can accept one of the two pairs of electrons present on the oxygen atom in water molecule. As a result of this, a hydroxonium ion (H_3O^+) is formed. (Fig 3.7)

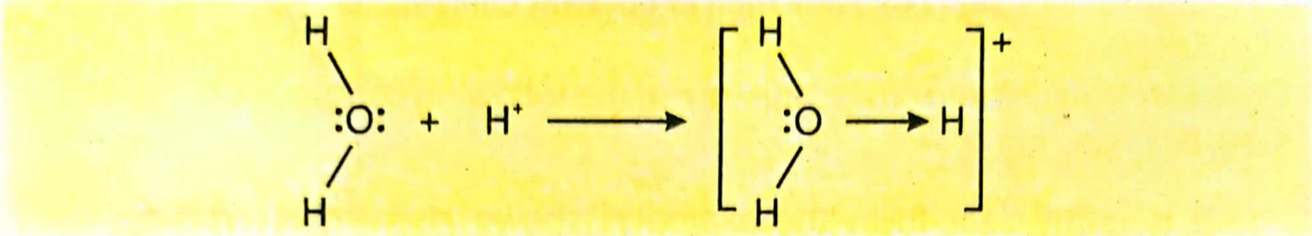


Fig (3.7): Formation of a Coordinate Covalent Bond Between H_2O and H^+

The positive charge covers whole of the hydroxonium ion. After the formation of hydroxonium ion, there does not remain any difference between a coordinate covalent bond and a covalent bond. All the three bonds of oxygen behave exactly alike.

Reaction Between ammonia and boron trifluoride

A reaction between ammonia (NH_3) and boron trifluoride (BF_3) is another example of the formation of a coordinate covalent bond. During the reaction, an electron pair from nitrogen of ammonia fills the partially empty outer shell of boron present in boron trifluorinde Fig (3.8).

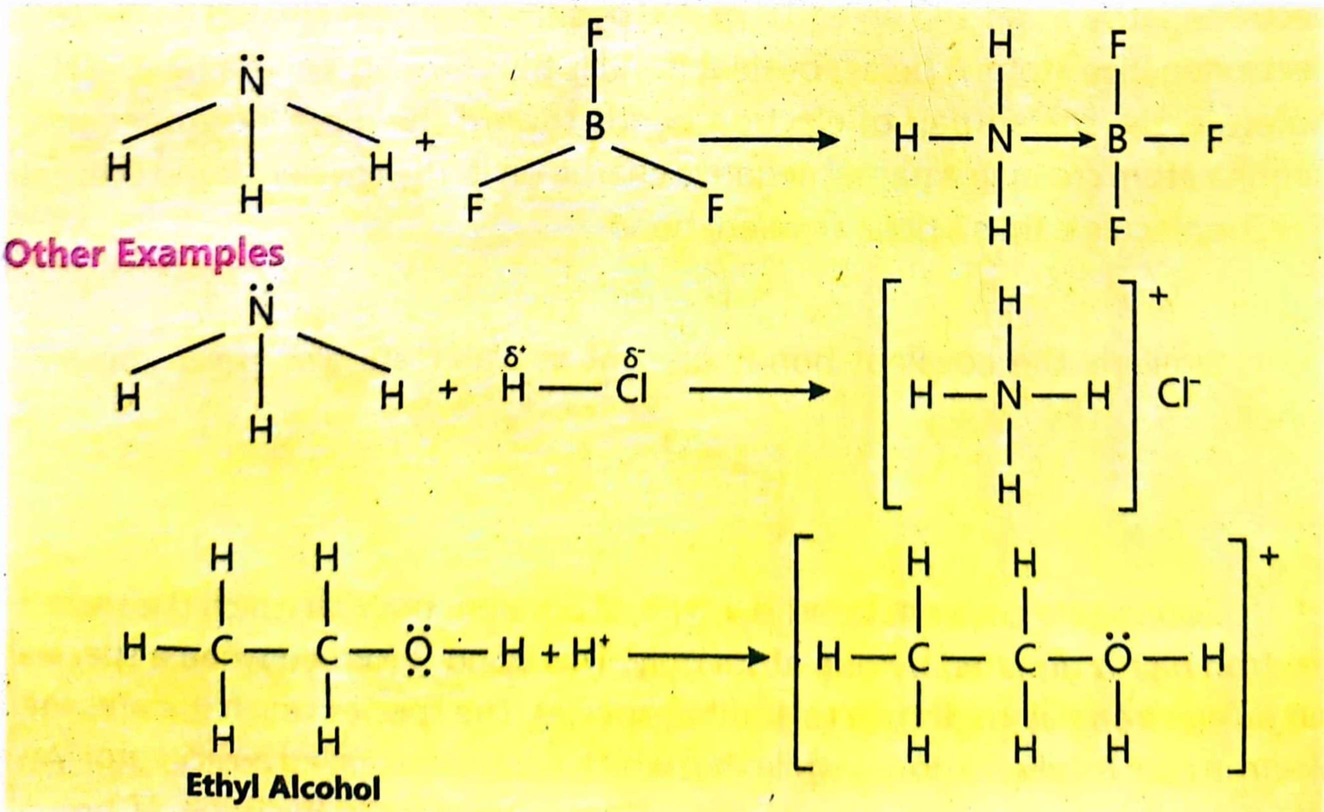


Fig (3.8): Formation of Boron trifluorinde ammonia, ammonium chloride and protonated ethyl alcohol

In the above example, a coordinate covalent bond in ammonium chloride links nitrogen of ammonia and the proton. The positive charge is spread all over ammonium ion. All the four bonds between nitrogen and hydrogen in ammonium ion behave exactly alike. This proves the point that the difference between a covalent bond and a coordinate covalent bond lies in the way they are formed. Once such bonds are formed, there does not remain any difference. When a coordinate covalent bond breaks both the bonded electrons go to the donor species.

Exercise

1. Draw the pictures of coordinate covalent bond formed between:
   (a) BF_3 and $AlCl_3$
   (b) $CH_3O\,CH_3$ and H^+
2. Which compound is not able to form a coordinate covalent bond?

3.3 Metallic Bond

The characteristics shown by metals are very different from those of ionic and covalent compounds. This suggests the presence of different types of binding forces among the metallic atoms.

Properties of Metals

1. Metals usually show metallic lustre.
2. Metals usually have high melting and boiling points.
3. Metals are good conductors of heat and electricity.
4. Metals are usually hard and heavy.
5. Metals can be made into different shapes by applying pressure.

These characteristics of metals can be explained if we know the nature of binding forces present between their atoms.

Usually metals have low values of ionization energy. Their atoms can therefore, lose their outer electron or electrons easily. In other words, the nuclei of metallic atoms cannot hold their outer electrons firmly. For example, in sodium metal, each sodium atom is surrounded by eight other sodium atoms. The outer electrons of these atoms move freely between the vacant spaces present between atoms because of the loose linkage they have with their nuclei. No electron remains attached with any particular nucleus. Instead, all the electrons, at the same time, get attached with all the nuclei. When all the atoms attract all the electrons collectively, obviously they will be bound together. A metal will appear to have a sea of electrons in which all the nuclei of atoms are

submerged. A metallic bond, is therefore a type of chemical bond which has positively charged ions bound together by the mobile electrons. Fig (3.9)

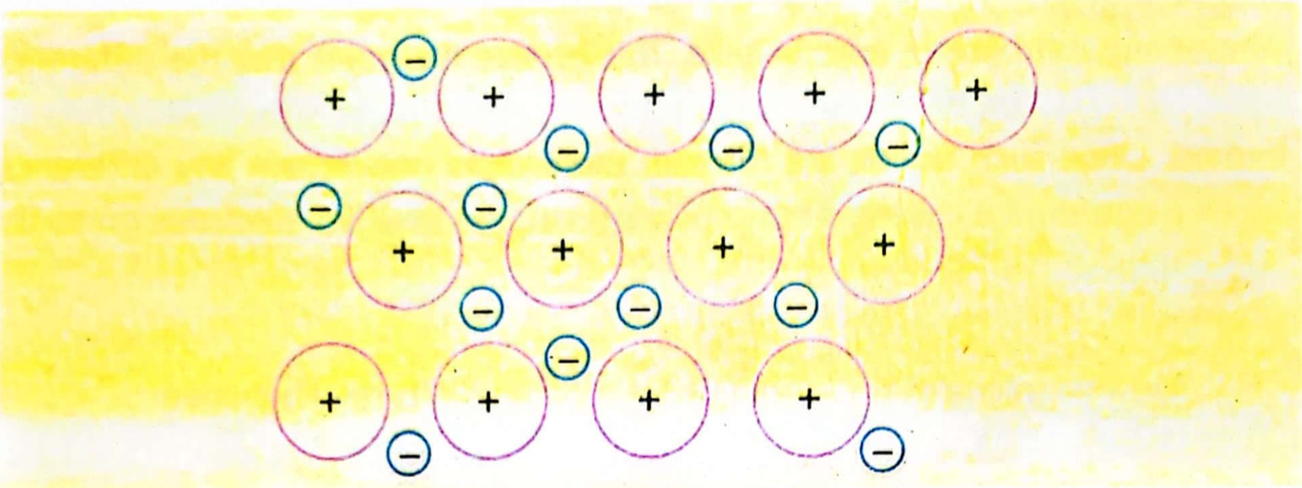


Fig (3.9): Metallic Bond in Sodium Metal

The strength of a metallic bond depends upon two factors: the number of positive charges present on the positive ions and the number of mobile electrons set free by each atom. In sodium metal, for example, each sodium atom sets free only one electron. The metallic bond in sodium metal is, therefore, not very strong. In magnesium metal, each magnesium atom releases two electrons to acquire two positive charges. The metallic bond in magnesium metal will evidently be stronger than that in sodium metal. This explains why the magnesium metal melts at a higher temperature than sodium metal. Fig (3.10)

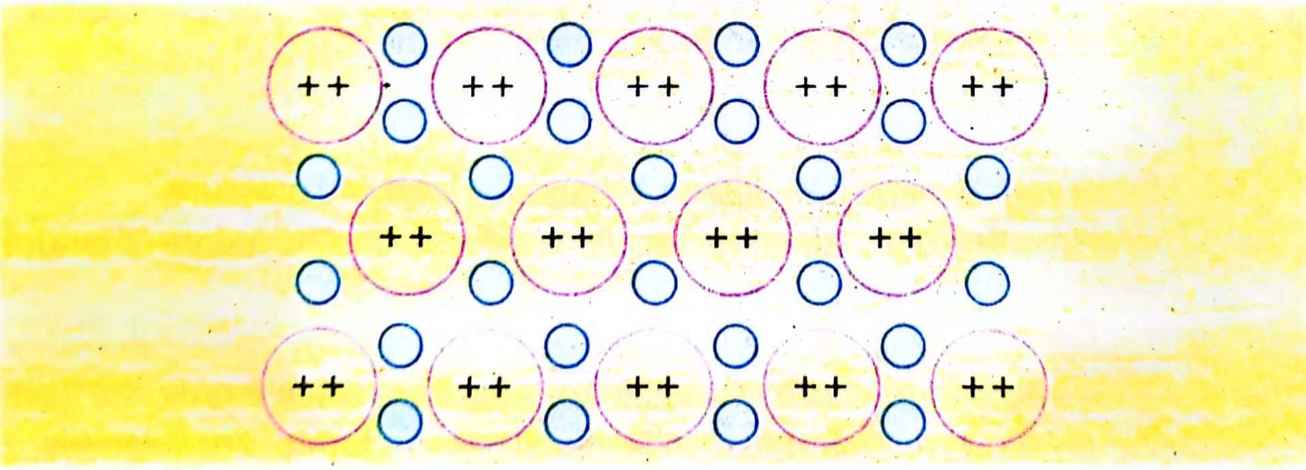


Fig (3.10): Metallic Bond in Magnesim Metal

The presence of freely moving electrons in metals makes them good conductor of heat and electricity. Moreover, in metals, the atoms are strongly held and arranged in the form of rows one above the other. This arrangement makes them hard and heavy. When pressure is applied on the metals, the upper rows of atoms slip pass the lower rows. As a result, their shapes are changed. Metals can, therefore, be easily drawn into wires and sheets. Fig (3.11)

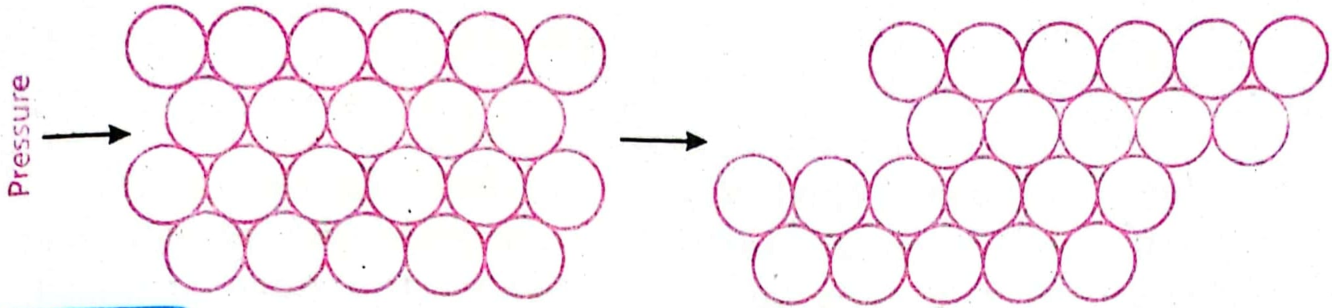


Fig (3.11) Metal under Pressure

Exercise

What type of atoms form metallic bond?
Give a comparison of metallic bond with an ionic bond.

Interesting Information!

Metals are extensively used in many industries. They are used in machinery, automobiles, railways, air crafts, rockets, in construction industry, in electronics industry, in jewellery, in electric wires and many more.

3.4 Electropositive Character of Metals

Metals generally have a tendency to lose electrons to form positive ions called cations. This property is called the electropositive character of metals. This property is also related to the reactivity of the metals. Metals which lose electron or electrons easily are considered more reactive. For example, alkali metals (Na, K) are highly electropositive elements and thus they undergo reactions very easily. Sodium and potassium react vigorously with water and halogens to give their respective hydroxides and halides. They also react with acids to give salts and hydrogen.

Alkaline earth metals (Mg, Ca), on the other hand, lose their outer electrons less easily and thus they are less electropositive than alkali metals. Their reactions towards water and halogens are also less vigorous.

Aluminum is also highly electropositive metal. It reacts readily with mineral acids to form salts and hydrogen.

3.5 Electronegative character of Non-metals

Non-metals have an affinity towards electrons. They tend to gain electrons and become negatively charged ions called anions. They are therefore, named as electronegative elements. Fluorine is the most electronegative element in the periodic table followed by oxygen, nitrogen and chlorine. Non-metals readily react with metals forming ionic bonds. Non-metals also combine with other non-metals to form a wide variety of molecular substances.

3.6 Compare the properties of ionic and covalent compounds.

Ionic Compounds.

1. In ionic compounds oppositely charged ions are properly arranged to give a crystalline structure. As a whole the compound is neutral. There exists a strong electrostatic force between their ions.
2. Ionic compounds are usually solids having high melting and boiling points. The melting point of sodium chloride is 801°C because it is difficult to break the strong electrostatic forces of attraction between the oppositely charged ions.
3. Ionic compounds are generally soluble in polar solvents like water.
4. They are usually good conductor of electricity in molten state or in solution form. Their conductance is due to the presence of free ions in these forces.

Covalent Compounds.

1. Covalent compounds mostly exist as discrete neutral molecules. There exists a strong electrostatic attraction between the nuclei and the shared electrons.
2. Covalent compounds are made of two or more non-metals. Lower molecular mass covalent compounds are gases or low boiling liquids. High molecular mass covalent compounds exist as solids. Generally, they have lower melting and boiling points.
3. They are usually insoluble in water but soluble in non-polar solvents like ether, benzene and acetone.
4. They are usually bad conductor of electricity.

3.7 Intermolecular Forces of Attraction

The forces of attraction which are present between the molecules of elements and compounds are named as intermolecular forces of attraction. These attractive forces are generally very weak as compared to the bonding forces present between the atoms of substances. Among the three states of matter, these forces are the weakest among the molecules of the gases and the strongest among the molecules of solids.

The intermolecular forces of attraction are of many type: some are weak and other are relatively strong. They affect the physical properties of the substances. The melting and boiling points of substances depend on the strength of these forces. The stronger the forces among the molecules of a liquid the higher is its boiling point and vice versa. Similarly stronger the intermolecular

forces the higher will be the melting point of a solid.

We shall explain here two type of such forces.

1. Dipole – Dipole Forces of Attraction

These attractive forces are present between the molecules of a polar compound like HCl. Hydrogen and chlorine attract the shared pair of electrons between them with different force. This force of attraction of an atom is called its electronegativity. Since the electronegativity of chlorine is greater than that of hydrogen it attracts the shared pair of electrons with greater force. As a result the bond between hydrogen and chlorine becomes polar as shown in the following:

$$H^{\delta+} \,:\, Cl^{\delta-}$$

Due to these partial charges the molecules of HCl start attracting each other. These forces of attraction are called dipole-dipole forces. (Fig 3.12)

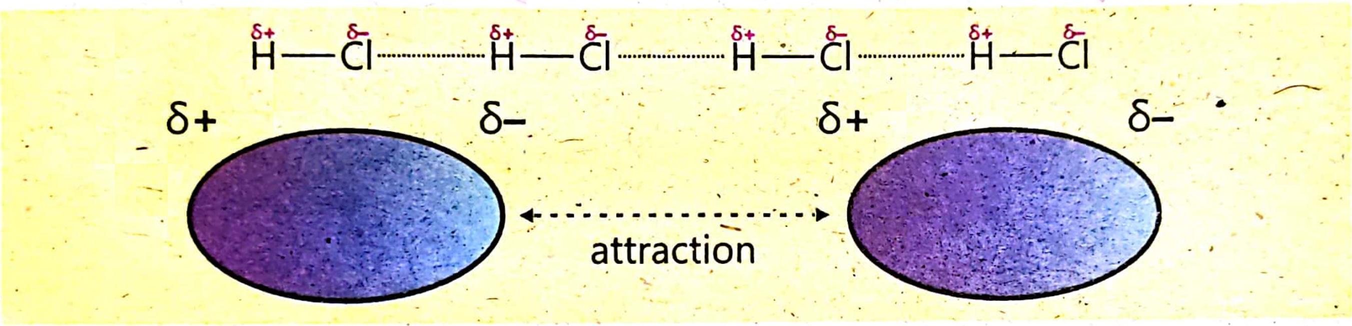


Fig (3.12): Dipole-Dipole Attraction

The compounds which have this type of attractive forces will show relatively higher melting and boiling points.

2. Hydrogen Bonding

Hydrogen bonding is a special case of dipole-dipole attractive forces. When hydrogen is covalent bonded to highly electronegative elements like F, O or N then the large difference of electronegativity values will make the covalent bond highly polar. As a result strong dipole-dipole attractions are observed among the molecules. For example, in H_2O, The O—H bonds are highly polar. Due to this, strong attractive forces are developed between water molecules as shown in the Fig (3.13).

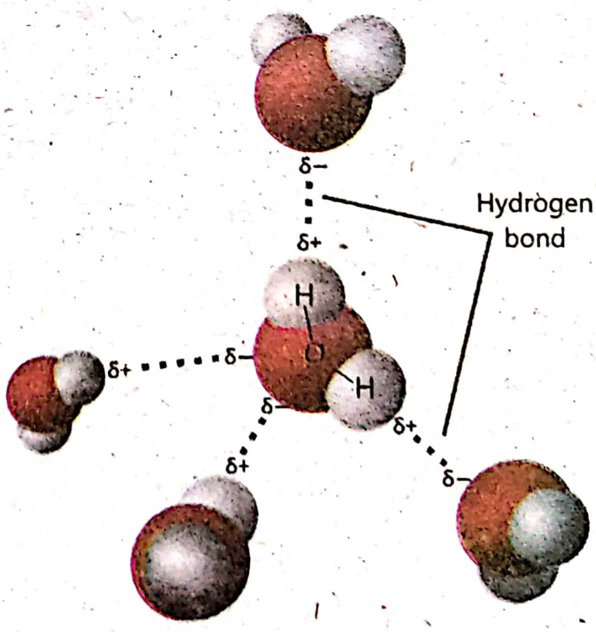


Fig (3.13): Hydrogen bond

This attractive force present between the molecules of water is called **Hydrogen Bonding**.

The strength of the hydrogen bonds causes water to have relatively higher melting and boiling points as compared to compounds like H_2S and NH_3.

3.8 Nature of Bonding and Properties

In ionic compounds, the oppositely charged ions are held together by the strong electrostatic force of attraction in the form of a crystal lattice. Since the ions are rigid in ionic compounds, such compounds therefore exist in the form of very stable solids with significantly high melting points. Since ions are spherical and oppositely charged they can surround each other from all the sides, ionic bonds are non-directional. This arrangement of ions is called crystal lattice.

If an external force is applied on the crystal lattice, it breaks easily. It shows that ionic solids are highly brittle. In the solid form, ionic compounds do not conduct electricity because ions are tightly held and cannot move. However, in the molten state, the ions get free and start conducting electricity. Ionic solids are also generally soluble in water. Water not only breaks the electrostatic force of attraction but also hydrates the resulting free ions Fig (3.14). In the process of hydration water molecules surround and interact with ions or molecules.

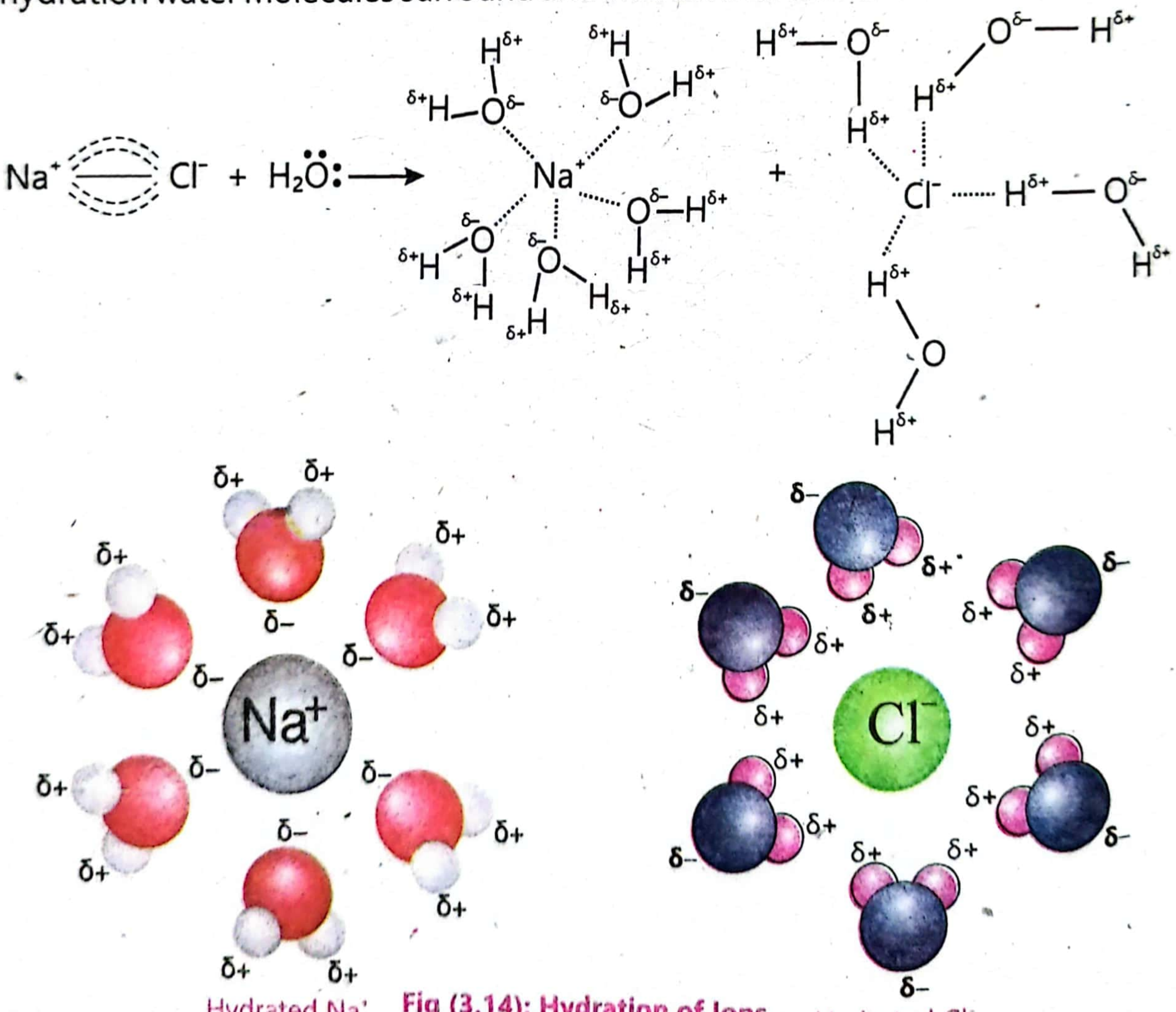


Fig (3.14): Hydration of Ions

Ionic compounds in an aqueous solution also conduct electricity because the free ions can now move towards their respective electrodes. Ionic compounds generally react in an aqueous solution. When we mix two solutions of ionic compounds, the positive ions of one compound may react with the negative ions of the other to form a new compound.

$$NaCl_{(s)} + AgNO_{3\,(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)} + Ag^+_{(aq)} + NO^-_{3(aq)}$$

$$\downarrow$$

$$AgCl_{(s)} + Na^+_{(aq)} + NO^-_{3(aq)}$$

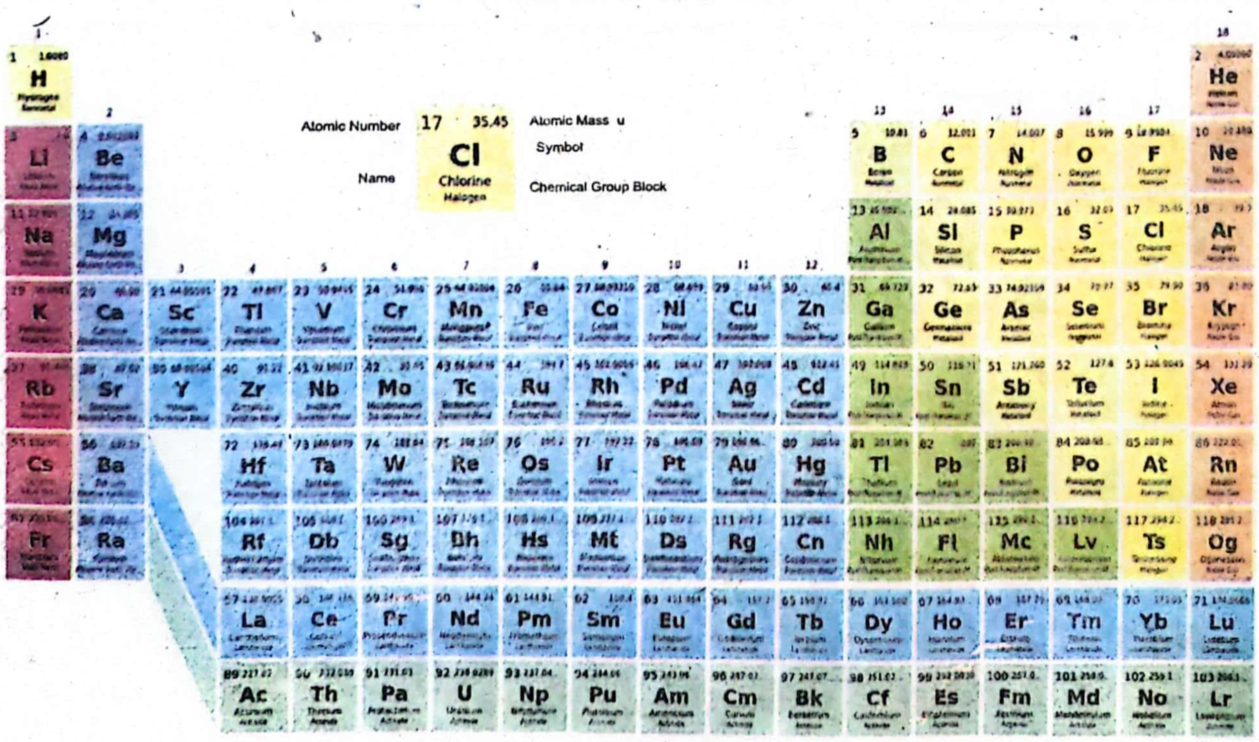
White precipitate of silver chloride comes out of the aqueous solution.

Interesting Information!

Conduction of ionic compounds in molten state and in form of an aqueous solution has been utilized to prepare many important elements and compounds. For example, electrolysis of molten sodium chloride gives us sodium metal and chlorine gas. Similarly electrolysis of aqueous sodium chloride gives sodium hydroxide and chlorine gas.

Diamonds, due to their exceptional hardness, are highly valued in industries. Diamond tipped glass cutters are used to make clean cuts in glass. Diamond-tipped drill bits are used to drill through hard rocks in mining operation.

Graphite is used in pencils, in polishes and to make crucibles. Graphite electrodes are used in battery cells and in electric arc furnaces to produce steel.



Elements and compounds in which atoms are covalent bonded, behave very differently from ionic compounds. Elements present at the right side of the periodic table exist as covalently bonded diatomic molecules except inert gases, for example, nitrogen(N_2), oxygen(O_2), fluorine(F_2) and chlorine(Cl_2). Due to very weak forces of attraction between their molecules, their densities and boiling points are very low. Bromine (Br_2) exists as volatile fuming liquid while elements like carbon, phosphorous and sulphur exist as covalent solids. All these solid elements exist both in amorphous and crystalline forms.

Coal is the amorphous form of carbon whereas diamond and graphite are its crystalline forms. Coal is used as a fuel in electricity generating plants. In diamond, each carbon atom is surrounded by four other carbon atoms linked together by strong covalent bonds. Due to this rigid structure, diamond is the hardest thing on this planet. It is used as a cutting, polishing and drilling tool.

Graphite consists of a layered structure, made of hexagonal rings of carbon. Since layers are not bonded strongly, they can slip past each other. Graphite is thus used as a lubricant in industry. Further, these layers in graphite have mobile electrons in between them. Graphite is a good conductor of electricity and it is also used as an electrode.

Binary covalent compounds generally exist as low temperature boiling gases except water. Methane (CH_4), ammonia (NH_3), hydrogen sulphide (H_2S), hydrogen chloride (HCl), nitrogen dioxide (NO_2), carbon dioxide (CO_2) and sulphur dioxide are all covalent compounds which are gases at room temperature.

Water and hydrogen fluoride, on the other hand, are liquids at room temperature. Liquid water has a high boiling point because strong intermolecular forces are present between its molecules. Covalent molecules like hydrogen chloride, sulphuric acid and nitric acid ionize completely in water behaving as very strong acids.

$$HCl_{(g)} \xrightleftharpoons{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$$

$$H_2SO_{4(l)} \xrightleftharpoons{H_2O} 2H^+_{(aq)} + SO_4^{2-}{}_{(aq)}.$$

Key Points

1. Atoms form bonds with other atoms to stabilize themselves by obeying duplet and octet rules.
2. The force of attraction which keeps the atoms together is called a chemical bond.
3. Bond which is formed by the transference of one or more electrons is called ionic bond.
4. A covalent bond is formed by the mutual sharing of electrons between atoms. A covalent bond may be single, double or triple.
5. When a covalent bond is formed between two identical atoms it is called a non-polar covalent bond.
6. When a covalent bond is formed between two non-identical atoms, which have different electronegative values, it is called a polar covalent bond.
7. Intermolecular forces are of two types; dipole-dipole attraction and hydrogen bonding.
8. When an electron pair is provided by an atom or ion or molecule to form a bond it is called a coordinate covalent bond.
9. Ionic solids are crystalline compounds with high melting and boiling points. They are generally soluble in aqueous solution.
10. Lower molecular mass covalent compounds are gases or low boiling liquids. Higher molecular mass covalent compounds exist as solids. They are bad conductors of electricity and are soluble in organic solvents.
11. Properties of ionic and covalent compounds are adequately explained on the basis of the type of attractive forces present between them.

Exercise

1. Tick (✓) the correct answer.

i. **When molten copper and molten zinc are mixed together, they give rise to a new substance called brass. Predict what type of bond is formed between copper and zinc.**

(a) Coordinate covalent bond (b) Ionic bond
(c) Metallic bond (d) Covalent bond

(ii) **Which element is capable of forming all the three types of bonds; covalent, coordinate covalent and ionic?**

(a) Carbon (b) Oxygen
(c) Magnesium (d) Silicon

(iii) Why is H_2O a liquid while H_2S is a gas?

(a) Because in water, the atomic size of oxygen is smaller than that of sulphur

(b) Because water is a polar compound and there exists strong forces of attraction between its molecules

(c) Because H_2O molecule is lighter than H_2S

(d) Because water can easily freeze into ice

(iv) Which of the following bonds is expected to be the weakest?

(a) C−C (b) Cl−Cl

(c) O−O (d) F−F

(v) Which form of carbon is used as a lubricant?

(a) Coal (b) Diamond

(c) Graphite (d) Charcoal

(vi) Keeping in view the intermolecular forces of attraction, indicate which compound has the highest boiling point.

(a) H_2O (b) H_2S

(c) HF (d) NH_3

(vii) Which metal has the lowest melting point?

(a) Li (b) Na

(c) K (d) Rb

(viii) Which ionic compound has the highest melting point?

(a) NaCl (b) KCl

(c) LiCl (d) RbCl

(ix) Which compound contains both covalent and ionic bonds?

(a) $MgCl_2$ (b) NH_4Cl

(c) CaO (d) PCl_3

(x) Which among the following has a double covalent bond?

(a) Ethane (b) Methane

(c) Ethylene (d) Acetylene

2. Questions for Short Answers

i. What type of elements lose their outer electron easily and what type of elements gain electron easily?

ii. Why do lower molecular mass covalent compounds exist as gases or low boiling liquids.

iii. Give one example of an element which exists as a crystalline solid and it has covalent bonds between its atoms.

iv. Which property of metals makes them malleable and ductile?
v. Is cordinate covalent bond a strong bond?
vi. Write down dot and cross formula of HNO_3.

3. Constructed Response Questions

i. Why HF is a liquid while HCl is a gas?
ii. Why covalent compounds are generally not soluble in water?
iii. How do metals conduct heat?
iv. How many oxides does nitrogen form. Write down the formulae of oxides?
v. What will happen if NaBr is treated with $AgNO_3$ in water?
vi. Why does iodine exist as a solid while Cl_2 exist as a gas?

4. Descriptive Questions

i. Explain the formation of an ionic bond and a covalent bond.
ii. How do ions arange themselve to form NaCl crystal.
iii. Explain the properties of metals keeping in view the nature of metellic bond.
iv. Compare the properties of ionic and covalent compounds.
v. How will you explain the electrical conductivity of graphite crystals?
vi. Why are metals usually hard and heavy?

5. Investigative Questions

i. The formula of $AlCl_3$ in vapour phase is Al_2Cl_6 which means it exists as a dimer. Explain the bonding between its two molecules?
ii. Explain the structure of sand (SiO_2).

Chapter 4

Stoichiometry

Student Learning Outcomes

After studying this chapter, students will be able to:

- State the formulae of common elements and compounds.
- Define molecular formula of a compound as the number and type of different atoms in one molecule
- Define empirical formula of a compound as the simplest whole number ratio of different atoms in a Molecule.
- Deduce the formula and name of a binary ionic compounds from ions given relevant information
- Deduce the formula of a molecular substance from the given structure of molecules.
- Use the relationship amount of substance= mass/ molar mass to calculate number of moles, mass, molar mass, relative mass (atomic/molecular/formula) and number of particles
- Define mole as amount of substance containing Avogadro's number 6.02×10^{23}) of particles
- Explain the relationship between a mole and Avogadro's constant.
- Construct chemical equations and ionic equations to show reactants forming products, including state symbols.
- Deduce the symbol equation with state symbols for a chemical reaction given relevant information.

Stoichiometry is an important concept in chemistry which helps us to calculate the amounts of reactants and products by using a balanced chemical equation. It is based on the law of conservation of mass which states that matter can neither be created nor destroyed. Therefore, the total mass of all the reactants must be equal to the total mass of all the products. The stoichiometric coefficient used to balance a chemical equation provides the mole ratio between reactants and products.

Stoichiometry is used in industry quite often to determine the amount of raw materials required to produce the desired amount of the products.

4.1 Chemical formula

Elements exist in different forms in this world. There are elements which exist in the form of aggregate of atoms. These elements are represented by their symbols alone, for example, Na, Ca, C, Fe, etc. On the other hand, elements like O_2, N_2, H_2 exist as discrete molecules in which their atoms are chemically bonded to each other. In ozone, three atoms of oxygen are bonded to each other, so its chemical formula is O_3.

Similar to elements, chemical compounds also exist in different forms. Common salt i.e. sodium chloride exists in the form of ions which are bonded together in the form of a crystal. Since ratio between its ions is 1:1, sodium chloride is represented by a formula unit NaCl. Similarly, the other ionic compounds are represented by their formula units which show the minimum ratio present between their ions. Examples are $CaCl_2$, Kbr and $BaCl_2$, etc.

Covalent compounds generally exist in discrete molecules in which atoms are bonded together. For example, water exists as molecules which are represented by the chemical formula H_2O. It shows that in one molecule of water two atoms of hydrogen are bonded to one atom of oxygen. Similarly, chemical compound, ammonia is represented by NH_3 and methane gas is represented by CH_4. A chemical compound is thus, represented by a chemical formula which is called the molecular formula of that compound and which shows all the types of atoms bonded together in one molecule of that compound. Examples of covalent compounds are HCl, HF, H_2S, PH_3, H_2O_2, H_2SO_4, CO_2, CO, C_6H_6, etc.

Exercise

How would you differentiate between the chemical formula of an element and that of a compound? Give examples. Write down the names of ionic and covalent compounds whose formulas have been given in this article.

4.2 Empirical Formula

Empirical Formula of a compound shows the simplest whole number ratio of atoms present in that compound. All the ionic compounds are represented by their empirical formulas. These formulas show the simplest ratio present between their ions. The empirical formula of calcium fluoride is CaF_2 which shows the ratio present between calcium and fluoride ions in its crystal.

For covalent chemical compounds, which exist as molecules, the empirical formulas may be different from their molecular formulas. For example, hydrogen

peroxide is represented by its molecular formula H_2O_2; its empirical formula will be HO. Similarly, a benzene molecule has C_6H_6 as its molecular formula; so its empirical formula will be CH. For water, the molecular formula and the empirical formula are both the same i.e. H_2O because there does not exist any minimum ratio between hydrogen and oxygen atoms.

Since an empirical formula does not tell us the actual number of atoms present in that compound, rather it represents the simplest ratio between atoms, it is possible that some compounds may have the same empirical formula. For example, both benzene (C_6H_6) and acetylene (C_2H_2) have the same empirical formula CH.

Exercise

Give two examples of the compounds which have same empirical and molecular formulas.

4.3 Chemical Formulas of Binary Ionic Compounds

In order to write down the formula of an ionic compound, first identify the cations and anions and the number of charges present on them. Finally combine the two ions together to form an electrically neutral compound.

If you know the name of binary ionic compound, you can write its chemical formula. Start by writing symbol of cation with its charge. Then write the symbol of anion with its charge and find out how many of these ions are needed to give an electrically neutral compound. For example, write down the formula of lithium oxide. The symbol of lithium cation with its single positive charge is Li^+. The symbol of anion is O^{2-}.

Let us now apply crisscross method to write the formula. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are then dropped.

Cation Anion

Li^{1+} O^{2-}

Li O Li_2O

Write down the formula of Aluminium oxide.

Cation Al^{3+} Anion O^{2-} Al_2O_3

Write down the formula of Magnesium nitride.

Cation Mg^{2+} Anion N^{3-} Mg_3N_2

Table (4.1) shows some important atoms, their ions and the charges on the ions

Interesting Information!

The composition of all the chemical products we use in our lives, such as shampoos, perfumes, soaps and fertilizers are formed using stoichiometric calculations. Without stoichiometry the chemical industry does not exist.

Table (4.1) Atoms and their Cations and Anions

Atoms and their Cations with charges		Atoms and their Anions and Cations with charges	
Atom	**Charge**	**Atom**	**Charge**
H	H^{1+}	O	O^{2-}
Na	Na^{1+}	N	N^{3-}
Li	Li^{1+}	Cl	Cl^{1-}
K	K^{1+}	Br	Br^{1-}
Mg	Mg^{2+}	I	I^{1-}
Ca	Ca^{2+}	Cu	Cu^{1+}, Cu^{2+}
Ba	Ba^{2+}	Fe	Fe^{2+}, Fe^{3+}
Zn	Zn^{2+}	Sn	Sn^{2+}, Sn^{4+}
Al	Al^{3+}	Ni	Ni^{2+}

4.4 Chemical Formulas of Compounds

Molecular formula of a compound can be found out if we know its empirical formula. To calculate the empirical formula of a compound, you need to determine the simplest whole-number ratio of atoms in the compound. This can be done by using experimental data on the mass percent composition of the compound. Molecular formula is then calculated by the following relationship.

$$\text{Molecular formula} = n\,(\text{Empirical Formula})$$

$$\text{Where } n = \frac{\text{Molar Mass}}{\text{Empirical Formula mass}}$$

For example, the empirical formula of hydrogen peroxide is HO. Its molar mass is 34. Its molecular formula will then be

$$n = \frac{34}{17} = 2$$

$$\text{Molecular formula} = (HO)_2 = H_2O_2$$

If for a compound the value of n is one, then its molecular formula is the same as its empirical formula.

Exercise

Write down the names of three such compounds which have different empirical and molecular formulas.

Sample Problem:

Empirical formula of a compound is CH. Its molecular mass is 78 g mol^{-1}. Find out its molecular formula

Solution:

Empirical Formula = CH   Molecular mass = 78 g mol-1

Molecular formula = n(Empirical formula)

n = molar mass / empirical formula mass

$$= \frac{78}{13} = 6$$

Molecular Formula $= (CH)_6 = C_6H_6$

Exercise

1. The empirical formula of a compound is CH_2O. Its molar mass is 180g mol^{-1}. Determine its molecular formula.
2. The empirical formula of a compound is CH_2O. Its molar mass is 60g mol^{-1}. Determine its molecular formula.

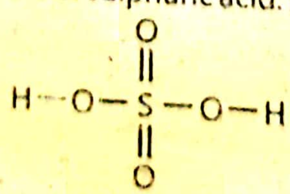
4.5 Deduce the molecular formula from the structural formula

In order to deduce the molecular formula from the structural formula the following steps are taken.

1. Write down the structural formula of the compound.
2. Count the number of atoms of each type in the structural formula.
3. Write the symbols of all the elements.
4. Write the total number of atoms of each kind as a subscript.
5. Remove the subscript 1.

Sample Problem:

Write down the molecular formula of sulphuric acid. Its structural formula is:

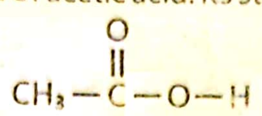


It has 2 H, 1 S and 4 O atoms.

Its molecular formula will be H_2SO_4.

Sample Problem:

Write down the molecular formula of acetic acid. Its structure formula is:

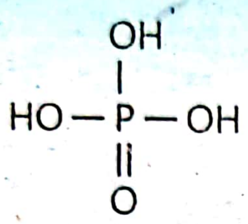


It has 2C, 4H, 2O atoms

Its molecular formula will be $C_2H_4O_2$

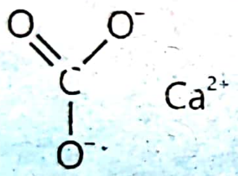
Exercise

1. Find out the molecular formula of phosphoric acid, its structural formula is:



2. Determine the molecular formula of n -- propyl alcohol. Its structural formula is $CH_3-CH_2-CH_2-OH$

3. Write down the formula of calcium carbonate. Its structural formula is:



4.6 Avogadro's Number (N_A)

In a chemical reaction, large number of atoms or molecules of reactants react to give the products. We would very much like to know the mass ratio in which the reactants react. For this purpose, we would also like to express these masses of reactants in grams. To achieve this objective, we need to transform the concepts of chemical formula and atomic mass units into such concepts which may lead us to know the masses of reacting elements and compounds in grams. Avogadro, an Italian scientist, helped us to achieve this objective in the following way.

Let us consider the following reaction in which two atoms of carbon react with a molecule of oxygen to produce two molecules of CO.

$$\underset{\text{Atom}}{2C} + \underset{\text{Molecule}}{O_2} \longrightarrow \underset{\text{Molecules}}{2CO}$$

Since it is not possible to account for the masses of individual atoms or molecules because these are very small particles, we increase the number of reacting species as written below.

$$2C \quad + \quad O_2 \longrightarrow 2CO \quad ....(1)$$

2 × 100 atoms	100 Molecules	2 × 100 Molecules
2 × 10000 atoms	10000 Molecules	2 × 10000 Molecules

Increasing the number of reacting atoms or molecules will not change the ratio in which these are reacting or are being formed.

Increasing the number of reacting species, however, has not solved the problem because this number is still very small. We should increase this number to such a value whereby it is convenient for us to calculate their masses.

Thus

$$2C \quad + \quad O_2 \longrightarrow 2CO \quad ....(2)$$

$2\times6.022\times10^{23}$ Atoms	6.022×10^{23} Molecules	$2\times6.022\times10^{23}$ Molecules

6.022×10^{23} is a huge number and we have selected this because

$$1.00\ g = 6.022 \times 10^{23}\ amu$$

Now the amounts of reactants and products in the forementioned equations can be written as follows.

$2 \times 6.022 \times 10^{23} \times (12.0\ amu) = 24.00$ g carbon atoms

$6.022 \times 10^{23} \times (32.0\ amu) = 32.00$ g oxygen molecules

$2 \times 6.022 \times 10^{23} \times (28.0\ amu) = 2 \times 28.00$ g carbon monoxide molecules

The mass ratio between the reactants and those of products will then become

$$C \quad + \quad O_2 \longrightarrow 2CO \quad ....(3)$$

24g	32g	56g

You must have realized that starting from a simple equation we have developed such ratio of masses of the reacting species which can conveniently be used in the laboratory.

According to the above-mentioned equation, 24 g of carbon contains $2 \times 6.022 \times 10^{23}$ atoms of carbon, 32 g of oxygen contains 6.02×10^{23} molecules of oxygen and 56 g of carbon monoxide contains $2 \times 6.022 \times 10^{23}$ of its molecules.

The number 6.022×10^{23} is called Avogadro's number after the name of an Italian chemist Amaedo Avogadro. Avogadro's number is the number of units in one mole of a substance. This number is represented as N_A.

4.7 The Mole and Molar Mass

Avogadro's number has an immense significance in Chemistry and the quantity of a substance containing Avogadro's number of particles (NA) is called **a Mole**. Mole is a number like a dozen or a gross. A dozen of oranges means 12

oranges. Similarly, a mole of a substance means its 6.022×10^{23} particles which can be atoms, molecules or ions. When we use the term mole of a substance, we must also refer to what type of particles are present in this substance. The following examples will help you to understand the concept clearly.

A mole of carbon atoms contains 6.022×10^{23} atoms and weighs 12 g.

A mole of oxygen molecules (O_2) contains 6.022×10^{23} molecules and weighs 32g. A mole of sodium chloride (NaCl) consists of 6.022×10^{23} of its formula units and its mass is 58.5 g.

The mass of one mole of a substance is called Molar mass. The molar mass of hydrogen atoms refers to the mass of one mole of hydrogen atoms and its value is 1.008 g. Similarly the molar mass of hydrogen molecules will be 2.016 g.

The chemical equation discussed in the previous article will now be understood in the following way.

$2C$	+	O_2	$\longrightarrow$	$2CO$
Two atoms		One molecule		Two molecules
Two moles of C		One mole of O_2		Two moles of CO
$2 \times 6.022 \times 10^{23}$ atoms		6.022×10^{23} molecules		$2 \times 6.022 \times 10^{23}$ molecules
Two moles		One mole		Two moles
24 g		32 g		56 g

Interesting Information!

Mole is important because atoms and molecules are so small that they can not be counted. The mole concept allows us to count atoms and molecules by weighing macroscopically small amounts of matter.

Exercise

Calculate the molar masses of the following compounds H_3PO_4, SiO_2, $C_{12}H_{22}O_{11}$, N_2O_4, $MgCO_3$

Sample Problems:

Determine the molar masses of the following compounds in g mol^{-1}.

(a) H_2SO_4 (Sulphuric acid) (b) $C_6H_{12}O_6$ (Glucose)

(a) H_2SO_4

Atomic mass of H = 1 Atomic mass of S = 32 Atomic mass of O = 16

Add the contribution of each atom

$2(1) + 1(32) + 4(16) = 98$ g mol^{-1}

(b) $C_6H_{12}O_6$

Atomic mass of C = 12 Atomic mass of H = 1 Atomic mass of O = 16

Add the contribution of each atom

$6(12) + 12(1) + 6(16) = 180$ g mol^{-1}

4.8 Chemical Equations and Chemical Reactions

To understand a chemical change and to study the different factors which control it, has always been a focal point in the efforts of chemists. To achieve these objectives, we should have an appropriate way of representing a chemical change. Fortunately, the chemists have developed a very suitable way of representing a chemical change in terms of symbols of elements and formulas of compounds. **Representing a chemical change in this way is called a chemical equation.** A chemical equation tells us the elements or compounds which are reacting and those which are being produced as a result of a chemical change. The reacting substances are called as **reactants** while those being produced are called **products.** It is customary to write the reactants on the left-hand side and the products on the right-hand side. An arrow head drawn from reactants to products separates the two. The following example will help you to understand these points clearly.

$$\underbrace{4Al_{(s)} + 3O_{2\,(g)}}_{\text{Reactants}} \longrightarrow \underbrace{2Al_2O_{3(s)}}_{\text{Products}}$$

The following points must be kept in mind while writing a chemical equation.

1. A chemical equation must obey the law of conservation of mass. This means that no atom should be destroyed or produced during a chemical change. The total number and the type of atoms must remain the same during a chemical change. Thus the total number and kind of atoms on both sides of the equation must be equal or the chemical equation must be balanced.
2. The formulas of elements and compounds must be written correctly.
3. A chemical equation must determine the correct mole ratio among the reactants, the products and between the reactants and the products.
4. A chemical equation must also point out the direction in which the change is proceeding.
5. It is a usual practice to show the normal **physical states** of reactants and products as a subscript. Solid, liquid and gas are symbolized as s, l and g respectively. Aqueous (aq) represents the solvated ion.

A chemical equation can, however, only be written if all the previous mentioned points are experimentally verified. For example, the nature of the products and their correct formulas must first be ascertained before writing a chemical equation. After the experimental verification of all the information about a chemical change, we should now be able to write a correct chemical

equation. The following equation has been written keeping in view all the points mentioned above.

$$Zn_{(s)} + H_2SO_{4\,(aq)} \longrightarrow ZnSO_{4\,(s)} + H_{2\,(g)}$$

According to this chemical equation, zinc reacts with sulphuric acid to give zinc sulphate and hydrogen gas. This chemical equation further tells us the mole ratio not only between reactants or products but also between reactants and products. According to the equation, one mole of zinc reacts with one mole of sulphuric acid to produce one mole of zinc sulphate and one mole of hydrogen gas.

Sometimes a chemical reaction moves both ways. In other words, the reactants react to give the products and the products, in turn, react to give the reactants back. Such reactions are called as **reversible reactions** and are indicated by ($\rightleftharpoons$) sign, e.g.:

$$N_{2\,(s)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)}$$

Reaction involving ions may also be shown in the form of chemical equation. Both $AgNO_3$ and NaCl are ionic compounds and are soluble in water. When they are mixed in water, they react to form products.

$$AgNO_{3\,(aq)} + NaCl_{(aq)} \longrightarrow AgCl_{(s)} + Na^+_{(aq)} + NO^-_{3\,(aq)}$$

AgCl being insoluble comes out of the aqueous solution in the form of a precipitate.

4.9 Calculations Based on Chemical Equation

A complete and balanced chemical equation tells us the mole ratio or molar mass ratio between the reactants and the products. With the help of this ratio, we can find out the molar masses of the products provided we know the molar masses of the reactants. Similarly the molar masses of the reactants can also be found out if we know the molar masses of the products.

For example, the following equation tells us that one mole (100 g) of calcium carbonate reacts with two moles (73 g) of hydrochloric acid to produce one mole (111 g) of $CaCl_2$, one mole (18 g) of water and one mole (44 g) of carbon dioxide.

$CaCO_{3\,(s)}$	+	$2HCl_{(aq)}$	$\longrightarrow$	$CaCl_{2\,(aq)}$	+	$H_2O_{(l)}$	+	$CO_{2\,(g)}$
1 mole		2 moles		1 mole		1 mole		1 mole
100 g		73 g		111 g		18 g		44 g

The total masses of the reactants are equal to the total masses of the products.

Example 1

25g of limestone ($CaCO_3$) reacts with an excess of hydrochloric acid according to the above equation. How much calcium chloride ($CaCl_2$) will be produced?

Solution Mass of $CaCO_3$ = 25 g   Molar mass of $CaCO_3$ = 100 g mol^{-1}

Mass of $CaCl_2$ product = ?   Molar mass of $CaCl_2$ = 111 g mol^{-1}

According to the equation

100 g of limestone reacts to produce calcium chloride = 111 g

1 g of limestone will react to produce calcium chloride $= \frac{111}{100}$ g

25 g of limestone will react to produce calcium chloride $= \frac{111}{100} \times 25$

$= 27.75$g

Example 2

1.80 moles of ethyl alcohol, when burnt in air completely, will utilize how many moles of oxygen gas? Also calculate the number of moles of CO_2 produced.

Solution No. of moles of ethyl alcohol = 1.80

No. of moles of oxygen needed = ?

The balanced chemical equation for the reaction will be

$$C_2H_5OH_{(l)} + 3O_{2(g)} \longrightarrow 2CO_{2(g)} + 3H_2O_{(g)}$$

According to this equation

One mole of ethyl alcohol, when completely burnt, needs oxygen = 3 moles

1.8 moles of ethyl alcohol, upon burning, will need oxygen $= \frac{3}{1} \times 1.8$

= 5.4 moles

ole of ethyl alcohol produces moles of CO_2 = 2.0

3 mol of ethyl alcohol will produce $= \frac{2}{1} \times 1.8$

= 3.6 moles

Example 3

Aluminium metal reacts with oxygen to produce aluminium oxide. How many grams of oxygen will be required to react completely with 0.3 moles of aluminium?

Solution Moles of Al = 0.3 Grams of O_2 used = ?

The balanced chemical equation for the reaction.

$$4Al_{(s)} + 3O_{2\,(g)} \longrightarrow 2Al_2O_{3\,(s)}$$

According to this equation

4 moles of aluminium need oxygen = 3.0 moles

1.0 mole of aluminium will need oxygen $= \frac{3}{4}$

0.3 moles of aluminium will need oxygen $= \frac{3}{4} \times 0.3$

= 0.225 moles

1 mole of oxygen (O_2) has molar mass = 32 g

0.225 mole of oxygen (O_2) will have molar mass = 32 × 0.225

= 7.2 g

Example 4

How many molecules of water will be produced if we react 5 g of hydrogen gas with excess of oxygen gas?

Solution Mass of H_2 used = 5g Molecules of water produced = ?

The balanced chemical equation for the reaction:

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$$

$2H_{2(g)}$	$O_{2(g)}$	$2H_2O_{(l)}$
2 moles	1 mole	2 moles
4g	32 g	36 g

According to this equation

4 g of hydrogen produce H_2O = 36 g of H_2O

5 g of hydrogen will produce H_2O $= \frac{36}{4} \times 5$ = 45g of H_2O

18g (1 mole) of H_2O contain molecules $= 6.022 \times 10^{23}$

36g of H_2O contain molecules $= 6.022 \times 10^{23} \times 2$

$= 12.04 \times 10^{23}$

45g of H_2O contain $= \frac{45}{36} \times 12.04 \times 10^{23}$

$= 1.505 \times 10^{24}$ molecules

Key Points

1. Molecular formula of an element or a compound shows the actual number of atoms present in the molecule of the element or a compound.
2. The formula of a compound which gives the minimum ratio present between its atoms is called its Empirical formula. All the ionic compounds and some of the covalent compounds are represented by their empirical formulas.
3. Chemical formula of a binary ionic compounds can be written if you know their names and the charges present on the ions.
4. Chemical formula of a compound is n times its empirical formula where n is the ratio of molar mass to empirical formula mass.
5. Avogadro's number has been calculated to know the mass ratio of reactants and products in a chemical reaction. The value of this number is 6.022×10^{23}.
6. The quantity of an element or a compound, which contains Avogadro's number of particles, is called a Mole and the mass of a substance present in it is called the Molar Mass.
7. A chemical equation tells the reacting and producing elements or compounds in a chemical reaction. It also tells the mole ratio between reactants or products and between reactants and products. A chemical equation must be balanced and should show the correct formulas of all the participating elements and compounds.
8. The mole ratio between reactants and products as shown by a chemical equation enables us to find out the mass ratio of these substances. A chemical equation is used to calculate the masses of the reactants as well as the products, which take part in a chemical reaction.

Exercise

1. Tick(✓) the correct answer.

(i) How many atoms are present in one gram of H_2O?

(a) 1002×10^{23} atoms (b) 6.022×10^{23} atoms
(c) 0.334×10^{23} atoms (d) 2.004×10^{23} atoms

(ii) Which is the correct formula of calcium phosphide?

(a) CaP (b) CaP_2
c) Ca_2P_3 (d) Ca_3P_2

(iii) How many atomic mass units (amu) are there in one gram?

(a) 1 amu (b) 10^{23} amu
(c) 6.022×10^{23} amu (d) 6.022×10^{22} amu

(iv) Structural formula of 2-hexene is $CH_3-CH=CH-(CH_2)_2-CH_3$. What will be its empirical formula?

(a) C_2H_2 (b) CH
(c) C_6H_{12} (d) CH_2

(v) How many moles are there in 25 g of H_2SO_4?

(a) 0.765 moles (b) 0.51 moles

(c) 0.255 moles (d) 0.4 moles

(vi) A necklace has 6g of diamonds in it. What are the number of carbon atoms in it?

(a) 6.02×10^{23} (b) 12.04×10^{23}

(c) 1.003×10^{23} (d) 3.01×10^{23}

(vii) What is the mass of Al in 204g of aluminium oxide, Al_2O_3.

(a) 26g (b) 27g (c) 54g (d) 108g

(viii) Which one of the following compounds will have the highest percentage of the mass of nitrogen?

(a) $CO(NH_2)_2$ (b) N_2H_4 (c) NH_3 (d) NH_2OH

(ix) When one mole of each of the following compounds is reacted with oxygen, which will produce the maximum amount of Co_2?

(a) Carbon (b) Diamond

(c) Ethane (C_2H_6) (d) Methane (Ch_4)

(x) What mass of 95% $CaCO_3$ will be required to neutralize 50 cm^3 of 0.5M HCl solution?

(a) 9.5g (b) 1.25g (c) 1.32g (d) 1.45g

2. Questions for Short Answers

i. Write down the chemical formula of barium nitride.

ii. Find out the molecular formula of a compound whose empirical formula is CH_2O and its molar mass is 180.

iii. How many molecules are present in 1.5 g H_2O?

iv. What is the difference between a mole and Avogadro's number?

v. Write down the chemical equation of the following reaction.

Copper + Sulphuric acid → Copper sulphate + Sulphur dioxide + Water

3. Constructed Response Questions

i. Different compounds will never have the same moleculer formula but they can have the same empirical formula. Explain

ii. Write down the chemical formulas of the following compounds.
Calcium phosphate, Aluminium nitride, Sodium acetate, Ammonium carbonate and Bismuth sulphate.

iii. Why does Avogadro's number have an immense importance in chemistry?

iv. When 8.657g of a compound were converted into elements, it gave 5.217g of carbon, 0.962g of hydrogen and 2.478g of oxygen. Calculate the percentage of each element present in this compound.

v. How can you calculate the masses of the products formed in a reversible reaction?

4. Descriptive Questions

i. Which conditions must be fulfilled before writing a chemical equation for a reaction?

ii. Explain the concepts of Avogadro's number and mole.

iii. How many grams of CO_2 will be produced when we react 10 g of CH_4 with excess of O_2 according to the following equation?

$$CH_{4\,(g)} + 2O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2H_2O_{(g)}$$

iv. How many moles of coal are needed to produce 10 moles of CO according to the following equation?

$$3C_{(s)} + O_{2(g)} + H_2O_{(l)} \longrightarrow H_{2(g)} + 3CO_{(g)}$$

v. How much SO_2 is needed in grams to produce 10 moles of sulphur?

$$2H_2S_{(g)} + SO_{2\,(g)} \longrightarrow 2H_2O_{(l)} + 3S_{(s)}$$

vi. How much ammonia is needed in grams to produce 1 kg of urea fertilizer?

$$2NH_{3\,(aq)} + CO_{2\,(aq)} \longrightarrow CO(NH_2)_{2\,(aq)} + H_2O_{(l)}$$

vii. Calculate the number of atoms in the following.

(a) 3g of H_2 (b) 3.4 moles of N_2 (c) 10g of $C_6H_{12}O_6$.

5. Investigative Questions

i. It is generally believed that drinking eight glasses of water every day is required to keep oneself hydrated especially in the summer. If a glass occupies 400 cm^3 of water on the average, how many moles of water are needed for a single adult?

ii. The chemical formula for sand is SiO_2 but the sand does not exist in the form of discrete molecules like H_2O. How has its formula been determined keeping in view its structure?

Chapter 5

Energetics

Student Learning Outcomes

After studying this chapter, students will be able to:

- Explain the idea of a chemical system and its connection with its surroundings influences energy transfer during a chemical reaction.
- Differentiate between exothermic and endothermic reactions by giving examples.
- State that thermal energy is called enthalpy change and recognize its sign as negative for exothermic and positive for endothermic reactions
- Define activation energy as the minimum energy that colliding particles must have for a successful collision.
- Explain that activation energy depends on reaction pathway which can be changed using catalysts or enzyme (detailed pathways not required)
- Draw, label and interpret reaction pathway diagram for exothermic and endothermic reaction which includes enthalpy change, activation energy (uncatalyzed and catalyzed), reactants and products
- Recognize that bond breaking is endothermic and bond making is exothermic processes.
- Explain that enthalpy change is sum of energies absorbed and released in bond breaking and bond forming
- Calculate enthalpy change of a reaction given bond energy values
- Explain how respiration (aerobic and anaerobic), an exothermic process, provides energy for biological systems and lipids as reserve stores of energy.

Energy exists in different forms which are often interconvertable. In chemical energetics we are mainly concerned with two forms of energy.

1. Chemical Energy: This energy is stored in a molecule in which atoms are bonded to each other.
2. Heat Energy: This form of energy is released when a bond is formed and absorbed when it is broken.

In energetics we study the energy changes that take place during a chemical reaction. These changes are caused by the making and breaking of bonds during a reaction. In most of the reactions the weak bonds of reactants break while in products new strong bonds are formed. Since energy is needed to break a bond while energy is evolved when a bond is formed, such reactions take place always with the evolution of heat. If a reaction is accompanied with the evolution of heat it is called an **exothermic reaction** and if heat is absorbed during a reaction it is called an **endothermic reaction**.

In energetics we not only encounter heat which comes out of a chemical reaction but also another quantity which is called enthalpy.

Enthalpy (H) or heat content, is defined as the total amount of thermal energy stored in a compound. The unit of its measurement is kJ mol^{-1}.

When the energy is absorbed during a reaction, the total enthalpy of the system increases. When energy is evolved during a reaction, the total enthalpy of the system decreases.

Chemical energetics is part of a broader field of chemistry called Thermodynamics. In energetics we study the flow of energy in a chemical reaction. Thermodynamics deals with the energy changes during chemical reactions and how these changes affect the properties of a chemical system.

Thomas Young was the first to use the word 'energy' to the field of physics in 1802.

5.1 System and Surounding

In Chemistry, any physical or chemical change under study may also be called a system. The chemical reaction includes reactants, products, catalyst, solvent and anything else which is important to study this reaction. Everything else which does not fall in this system is called the surrounding. For example, if you are boiling water in the beaker, the water molecules will be called a system while everything surrounding this like beaker, burner, etc. will be called the surrounding. When energy is transferred from surrounding to the system, the change is called endothermic and it has a positive sign. When energy is transferred from system to surrounding, the change is called exothermic and it carries a negative sign, Fig. (5.1).

Interesting Information!

Energy evolved during a chemical reaction is used in everyday life for cooking, heating, lighting, transportation and much more.

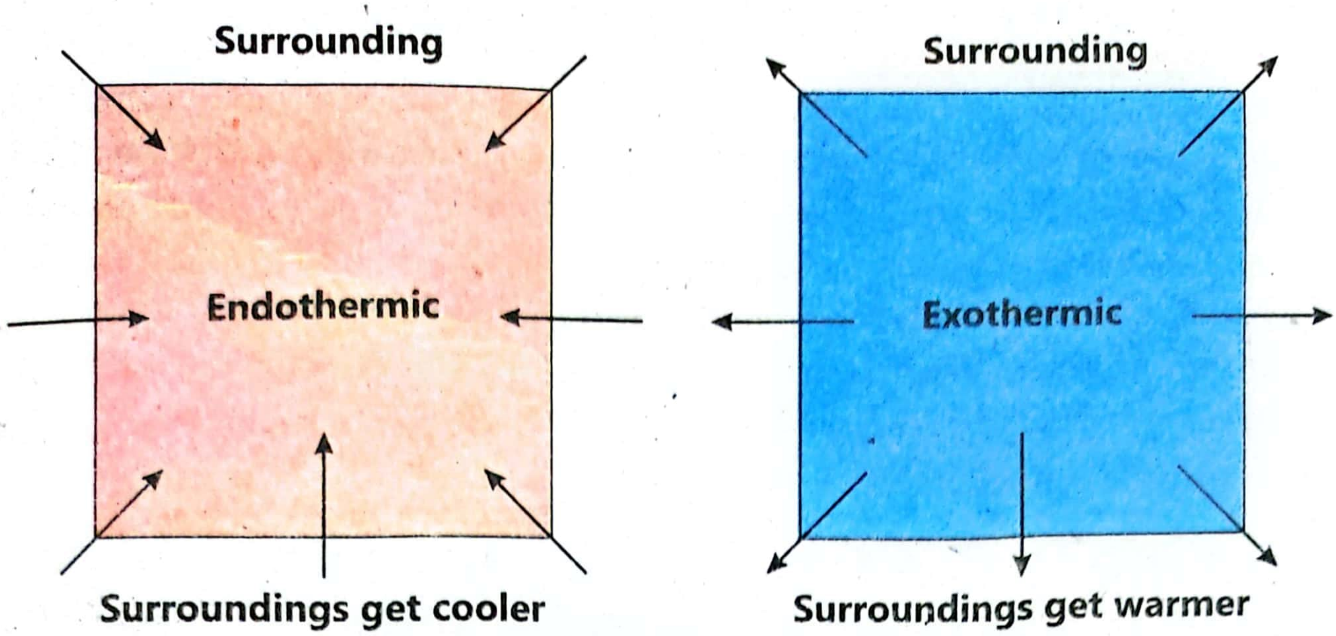


Fig (5.1) System and surrounding exchanging energy

Exercise

Does boiling water in a beaker endothermic or exothermic change? Which form of energy is being transferred in this system.

Interesting Information!

Enthalpy is important because it tells us how much heat is present in a system. Heat is important because we can extract useful work from it.

5.2 Enthalpy

The total amount of heat energy present in a molecule under standard conditions (0°C temperature and 760 mm pressure) is also called its heat content. Enthalpy is the measurement of energy in a thermodynamic system. The quantity of enthalpy is equal to the total heat content of a system. Enthalpy of a system is represented by (H) while the change in enthalpy which a system undergoes is represented by ΔH. The total enthalpy (H) of a system cannot be measured directly. However, the change in enthalpy ΔH brought about in a system can be measured comparatively easily.

In Chemistry, the standard enthalpy of reaction (ΔH°) is the enthalpy change when reactants in their standard states undergo reaction to produce products in their standard states. This quantity is called the standard enthalpy change or heat of reaction at constant pressure.

$$2\,CO_{(g)} + O_{2(g)} \longrightarrow 2\,CO_{2\,(g)} \qquad \Delta H^\circ = -566.0\ kJ$$

The reaction in this system is thus exothermic evolving 566 kJ of heat energy which is given to the surrounding.

How is enthalpy different from heat?

Heat is a form of energy that flows from hot body to a cold body because of a difference in temperature. We measure heat in joules. Heat is what we call the transfer of thermal energy. Contrary to this, enthalpy is an essential part of a system since it depends on the number of molecules present in that system, its chemical composition and its structure. Heat is not essential part of a system, it just comes and goes. When heat leaves or enters a system, it results in a change of enthalpy. At a constant pressure the enthalpy change is equal to heat evolved or absorbed.

Exercise

1. Can energy be transferred in a form other than heat during a chemical reaction?
2. Why it is not possible to calculate the enthalpy of a system?

5.3 Exothermic and Endothermic Reactions

A physical or a chemical change is almost always accompanied with either absorption or evolution of heat. Heat, which is evolved or absorbed during a chemical reaction, is called the heat of that reaction.

Chemical reactions in which heat energy is evolved are called exothermic reactions while those in which heat energy is absorbed are called endothermic reactions. Heat, which is evolved during an exothermic reaction, goes to the surrounding and the container in which such a reaction is being carried out, gets hot. Conversely, in an endothermic reaction, the absorption of heat from the surrounding will decrease the temperature of the container.

Example: Hydrogen gas and oxygen gas react to give liquid water in an exothermic reaction.

$$\underset{\text{Hydrogen}}{2H_{2(g)}} + \underset{\text{Oxygen}}{O_{2(g)}} \longrightarrow \underset{\text{Water}}{2H_2O_{(l)}} + 571.6\ \text{kJ}$$

571.6 kJ heat energy is evolved during this reaction. If the energy evolved is shown separately it is expressed as $\Delta H = -571.6$ kJ. The same amount of energy will be absorbed when the reaction will move in the backward direction i.e. water will decompose to give hydrogen and oxygen back.

Example: Carbon dioxide gas is produced when solid carbon burns in oxygen gas.

$$\underset{\text{Carbon}}{C_{(s)}} + \underset{\text{Oxygen}}{O_{2\,(g)}} \longrightarrow \underset{\text{Carbon dioxide}}{CO_{2\,(g)}} + 393.5\ kJ$$

It is also an exothermic reaction and 393.5 kJ heat energy is evolved during this reaction. When this reaction is carried out in the reverse direction, the same amount of energy i.e. 393.5 kJ will be absorbed. The enthalpy change (ΔH) for this reaction is -393.5 kJ mol^{-1}.

Both the reactions mentioned above are the examples of exothermic changes.

The following reactions represent endothermic changes.

$$\underset{\text{Hydrogen}}{H_{2\,(g)}} + \underset{\text{Iodine}}{I_{2\,(s)}} \longrightarrow \underset{\text{Hydrogen Iodide}}{2HI_{(g)}} - 53.08\ kJ$$

The enthalpy change for the reaction is $\Delta H = 53.08$ k.J

Hydrogen gas reacts with solid iodine only at high temperature and 53.08 kJ of heat energy is absorbed.

Formation of NO in air due to lightning in the clouds.

$$\underset{\text{Nitrogen}}{N_{2\,(g)}} + \underset{\text{oxygen}}{O_{2\,(g)}} \longrightarrow \underset{\text{Nitric Oxide}}{2NO_{(g)}} - 180.6\ kJ$$

The enthalpy change for the reaction is $\Delta H = 180.6$ kJ.

Interesting Information!

Heat evolved or absorbed during a reaction is used in self-heating or self-cooling packs. These packs contain reactants that undergo an exothermic or an endothermic reaction providing high or low temperature.

Our present-day living conditions depend heavily on the availability of energy in its various forms. Exothermic chemical reactions are extensively used to fulfill this requirement. In such reactions, chemical energy is converted into heat energy. We burn fuels like gas, oil and coal to cook food and for other heating purposes in our homes and industry. During this burning process called

combustion, substances present in fuels react with oxygen of the air to produce a certain amount of heat.

Foods such as fats and carbohydrates are important biological fuels. During metabolism, the chemical energy present in this food is converted to heat to keep us warm.

A large portion of electricity is produced at power stations by burning fuels such as natural gas, oil and coal. The heat which comes out from their combustion is used to produce steam at high pressure. This high pressure steam is then used to rotate turbines, which in turn generate electricity.

While driving a vehicle, it is the combustion of petrol or diesel that gives off energy and drives it forward. The one example of exothermic reactions people seem to enjoy the most is that of fireworks. Fireworks are the result of combustion reactions that yield heat, light and sound. Different metal salts alongwith oxidising agents produce a variety of colours when burnt, Fig. (5.2).

Fig (5.2) Fireworks

Exercise

1. Why the chemical reaction between sodium metal and water proceeds violently?
2. Is melting of ice an exothermic or endothermic change?
3. Can exothermic reaction be reversed?

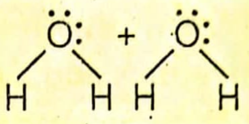
Let us now examine the reason why the chemical reactions are either exothermic or endothermic.

A chemical reaction mainly involves the processes which involve bond breaking and bond formation. In the following reaction, the chemical bonds between the atoms present in the molecules of H_2 and O_2 first break to give their atoms.

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$$

$$\underset{\text{Molecule}}{2H_{2(g)}} \longrightarrow \underset{\text{Atoms}}{4H_{(g)}} \qquad \underset{\text{Molecule}}{O_{2(g)}} \longrightarrow \underset{\text{Atoms}}{2O_{(g)}}$$

These atoms of hydrogen then form bonds with oxygen atoms to form two molecules of gaseous H_2O.

$$4H_{(g)} + 2O_{(g)} \longrightarrow$$ 

Breaking of bonds of H_2 and O_2 absorb energy (endothermic process) while making of bonds between H and O evolve energy (exothermic process). In this reaction, weaker bonds are broken i.e of H_2. Hence less energy is absorbed in the system. While the bonds which are formed in water molecule are stronger and thus greater energy is evolved. Hence, the energy which is evolved is more than the energy which is absorbed. The overall reaction is thus exothermic.

$$H_{2(g)} \longrightarrow 2H_{(g)} + 435\ kJ/mol \quad \text{(Endothermic)}$$

$$O_{2(g)} \longrightarrow 2O_{(g)} + 498\ kJ/mol \quad \text{(Endothermic)}$$

$$O + H \longrightarrow O-H - 484\ kJ/mol \quad \text{(Exothermic)}$$

Since two moles of H_2 take part in the reaction so total energy absorbed in the reaction

$$2H_{2(g)} \longrightarrow 4H_{(g)} + 870\ kJ$$

$$O_{2(g)} \longrightarrow 2O_{(g)} + 498\ kJ/mol$$

$$2H_{2(g)} + O_{2(g)} \longrightarrow 4H_{(g)} + 2O_{(g)} + 1368\ kJ$$

It means 1368 kJ energy is absorbed when 2 moles of H_2 and one mole of O_2 break their bonds to convert themselves into atoms.

Bond dissociation energy of H_2 is 435 kJ/mol while that of O_2 is 498kJ/mol. Bond formation energy of one O – H bond is 484 kJ/mol.

Interesting Information!

Nitrogen of the atmosphere reacts with oxygen to produce nitric oxide in the presence of lightning. This is because reaction is highly endothermic, so only lightning can provide enough energy for this reaction to take place.

Total energy evolved in the formation of 4 O – H bonds.

$$2\,O_{(g)} + 4H_{(g)} \longrightarrow 2\,H_2O_{(g)} - 1936\text{ kJ}$$

This is the energy evolved when two moles of water are formed from 4 moles of hydrogen atoms and 2 moles of oxygen atoms. Thus for the formation of one mole of water, the energy evolved will be 968 kJ mol^{-1}.

Hence the overall energy evolved in this reaction is = –1936 + 1368 = –568kJ for two moles of water.

$$2\,H_{2(g)} + O_{2(g)} \longrightarrow 2\,H_2O_{(g)} + 568\text{ kJ} \quad \text{or} \quad \Delta H = -568\text{ kJ}$$

The enthalpy change for the formation of two moles of gaseous water is thus –568kJ. So the enthalpy change for the formation of one mole of gaseous water will be

$$= -\frac{568\text{ kj}}{2\text{ mol}} = -284\text{ kJ mol}^{-1}$$

The experimental value of formation of gaseous water is -284.3 kJ which is quite close to this calculated value.

Sample Problem:

Calculate the enthalpy change of the following chemical reaction.

$$H_{2(g)} + I_{2(g)} \longrightarrow 2\,HI_{(g)}$$

Bond energies of H_2, I_2 and HI are 436,151, and -299 kJ mol^{-1} respectively.

Exercise

Calculate the enthalpy change for the formation of one mole of liquid water.

5.4 How does a Reaction take place?

A reaction takes place when the reactant molecules collide with each other to give a transition state. Let us study the following hypothetical reaction.

$$A_2 + B_2 \longrightarrow 2AB$$

Before mixing, the molecules of reactants A and B are in a state of random motion separately colliding with each other and with the walls of container. Kinetic energies possessed by these molecules are not the same. Majority of these molecules possess average kinetic energy but a few possess more than average energy while yet others possess less than average kinetic energy. The molecules which possess more than average kinetic energy may also be called excited molecules.

When the two reactant molecules are mixed together, all these molecules start colliding with each other. The collisions which result by colliding molecules having average or less than average kinetic energies may not be able to produce any result. But when the two excited molecules from both the reactants collide with each other they may be able to produce the transition state as shown in the following Fig (5.3).

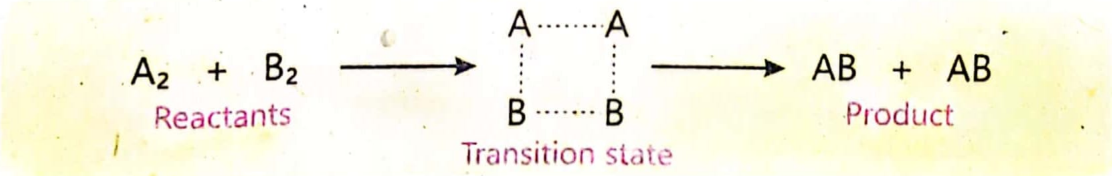


Fig (5.3) Formation of Transition state

The transitation state is shown at the peak of curve in Fig (5.4). After a very short period of time the transition state either returns to the reactants or to the products. The progress of the reaction can be shown in the form of the following energy profile diagram drawn between path of the reaction and the energy of the reactants and the products.

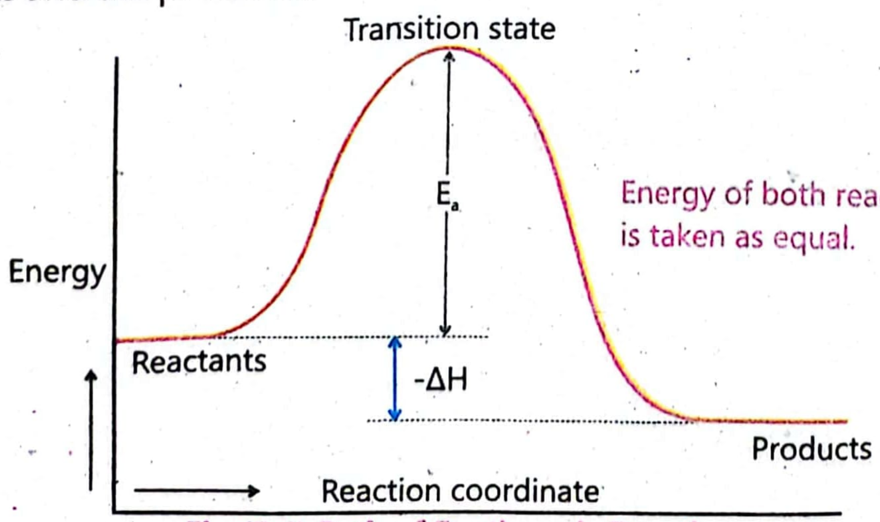


Fig (5.4) Path of Exothermic Reaction

Interesting Information!

Washing clothes at 60°C uses almost twice the energy as at 30°C wash. 90% of the energy used by the traditional electric bulb is wasted in producing heat.

The energy of the transition state is higher than that of reactants or products because the bonds between the reactant or product molecules are being cleaved progressively. The energy absorbed by the reactant or product molecules in order to be converted into the transition state is called the activation energy (E_a) of the reaction. The difference between the energy of reactant and that of the product comes out in the form of heat representing enthalpy (ΔH) of the reaction. This graph represents the path of an exothermic reaction Fig (5.4). A similar graph can be drawn for an endothermic reaction Fig (5.5).

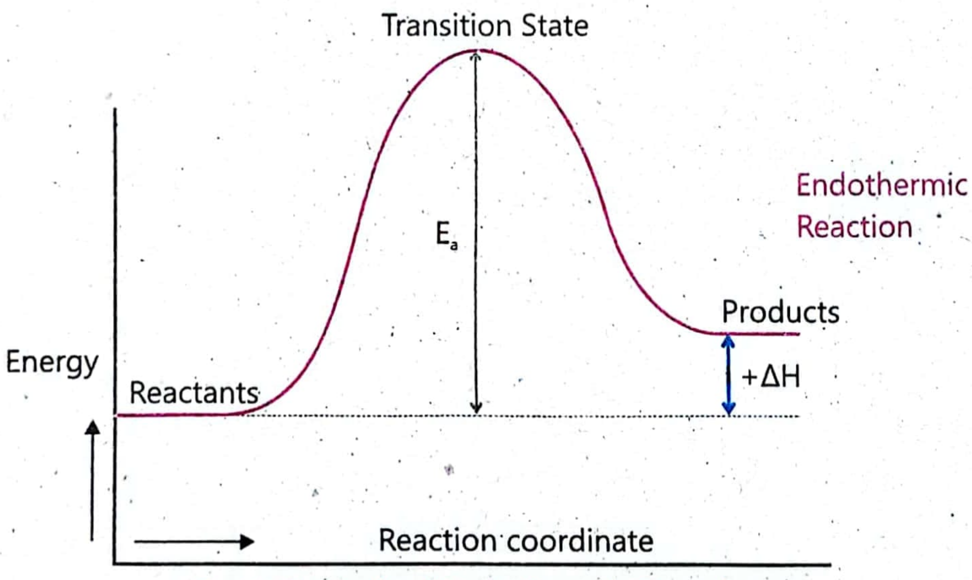


Fig (5.5) Path of Endothermic Reaction

Exercise

1. Are energy diagrams useful?
2. Draw an energy profile diagram for a hypothetical reaction which does not evolve or absorb heat.

An addition of the catalyst in a reaction increases the rate of reaction because it changes the path adopted by the reactants whereby the activation energy value of the reaction is substantially decreased. As a result, more reactants are now able to be converted into product molecules and hence the rate of reaction will increase Fig (5.6).

A catalyst is thus, defined as a substance that increases the rate of a chemical reaction without itself undergoing any permanent chemical change. **For example,** Ni acts as a catalyst in the hydrogenation of oil to give banaspati ghee. Platinum acts as a catalyst in the production of H_2SO_4. Chlorine acts as a catalyst promoting the breakdown of ozone.

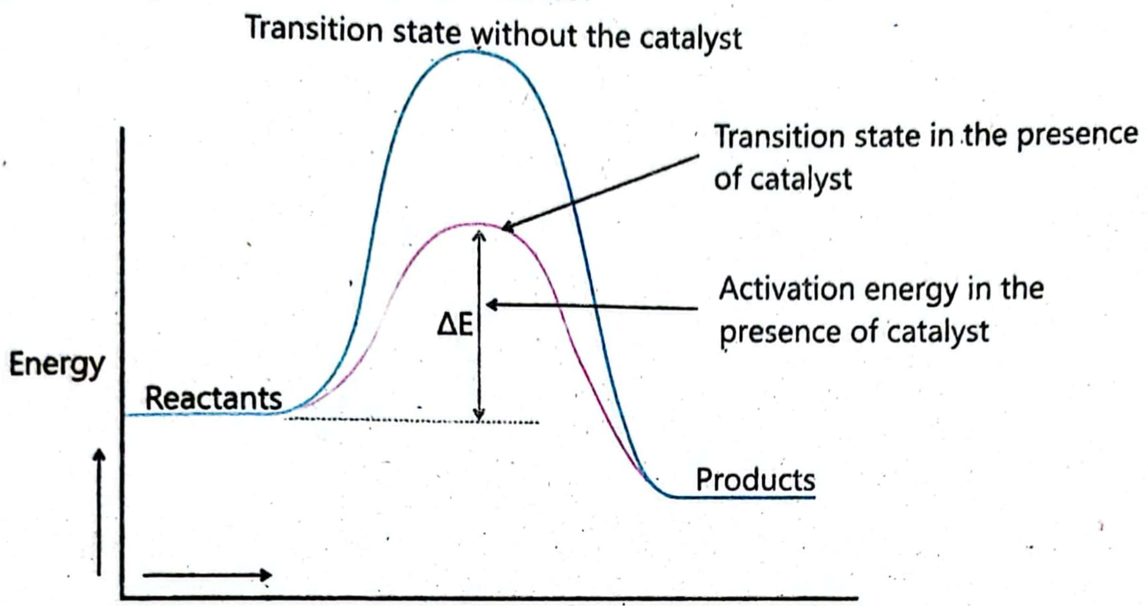


Fig (5.6) Path of Reaction in the absence and presence of a catalysist

5.5 Aerobic and Anaerobic Respiration

The process of respiration in human beings is a continuous process. During this process, we breathe in oxygen and breathe out carbon dioxide. Respiration also carries complex chemical reactions inside the human body. This process that occurs in the presence of oxygen is called aerobic respiration. Aerobic respiration is an exothermic process and involves the following reactions.

$$\underset{(C_6H_{12}O_6)}{\text{Glucose}_{(s)}} \xrightarrow{\text{in cytoplasm}} \underset{\text{Energy}}{\overset{(C_3H_4O_3)}{2\text{Pyruvate}} + } \xrightarrow[\text{in mitochondria}]{O_2} 6CO_{2(g)} + 6H_2O_{(l)} + \text{Energy}$$

Glycolysis

During glycolysis one molecule of glucose is split into two molecules of pyruvate. This process involves a series of reactions catalysed by enzymes, with a net production of 2 ATP (Adenosine Triphosphate). When cells of our body require energy for performing the metabolic activities, they use this ATP and break it down to get the required energy. The food we eat undergoes digestion in our body and the digested food molecules that are absorbed by the cells undergo oxidation to produce energy.

In certain organisms like bacteria and algae respiration occurs in the absence of oxygen and it is called anaerobic respiration. This is also an exothermic process and during this process glucose is converted into carbon dioxide and ethanol with the evolution of energy.

$$\text{Glucose}_{(s)} \xrightarrow{\text{in cytoplasm}} \text{Pyruvate} + \text{Energy} \xrightarrow[\text{Yeast}]{\text{Absence of oxygen}} C_2H_5OH_{(l)} + Co_{2(g)} + \text{Energy}$$

Lipids are a group of organic compounds which include fats, waxes, sterols, etc. Lipids serve as an energy reserve within our body. About half of the fuel our body needs comes from lipids. If you eat more food than you need in a day, the excess food is stored as lipids in adipose cells. In between meals and during exercise our body relies on this resource to provide energy. Glycogen is the primary storage form of glucose. It is stored in the liver and muscles.

Key Points

1. Any physical or chemical change under study is called a system. Everything else which does not fall in this system is called surrounding.
2. The total amount of heat energy present in a system at standard conditions is called its heat content. The quantity of enthalpy of a system is also called its total heat content.
3. The standard enthalpy of a reaction is the enthalpy change that occurs when reactants in their standard states undergo reaction to produce products in their standard states.
4. Chemical reactions which take place with the evolution of heat are called exothermic reactions while those which proceed with the absorption of heat are called endothermic reactions.
5. A chemical reaction always involves breaking and formation of chemical bonds.
6. When weaker bonds are broken while stronger bonds are formed, the reaction is overall exothermic and vice versa.
7. A reaction takes place when the reactant particles collide to give a transition state which then change into the products.
8. The energy needed by the reactant particles to change into transition state is called activation energy.
9. The difference between the energy of reactants and that of products is called enthalpy change of the reaction.
10. The reaction is exothermic if the energy of the products is less than that of reactants and endothermic if the energy of the products is more them that of reactants.
11. A catalyst increases the rate of reaction by decreasing its activation energy.
12. The process of respiration in human beings takes place in the presence of oxygen and it is called an aerobic respiration. In some organisms, the respiration occurs in the absence of oxygen which is called anaerobic respiration.

Exercise

1. Tick (✓) the correct answer.

(i) The following reaction is an exothermic reaction.

$$H_2 + Cl_2 \xrightarrow{\text{Sunlight}} 2HCl$$

From where does the energy come to break the bonds of H_2 and Cl_2?

(a) By collisions between the molecules

(b) From sunlight

(c) From the surrounding

(d) By collisions of the molecules with the walls of the container

(ii) Which of the following reactions has the least value of activation energy?

(a) $H_{2(g)} + \frac{1}{2} O_{2(g)} \longrightarrow H_2O_{(g)}$

(b) $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$

(c) $NaCl_{(aq)} + AgNO_{3(aq)} \longrightarrow AgCl_{(s)} + NaNO_{3(aq)}$

(d) $H_{2(g)} + I_{2(s)} \longrightarrow 2HI_{(g)}$

(iii) Formation of which hydrogen halide from the elements is an endothermic reaction?

(a) HCl (b) HF

(c) HBr (d) HI

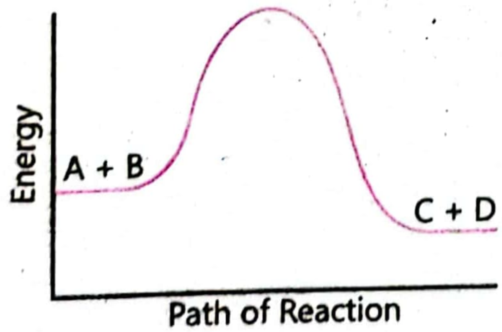
(iv) What are the products of anaerobic respiration?

(a) ATP + CO_2 + H_2O (b) CO_2 + H_2O

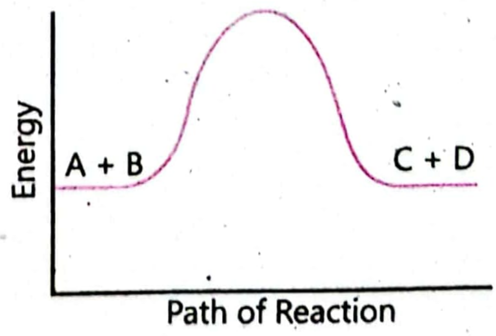
(c) ATP + Ethanol + H_2O (d) Ethanol + H_2O

(v) Which reaction do you expect to be a reversible reaction?

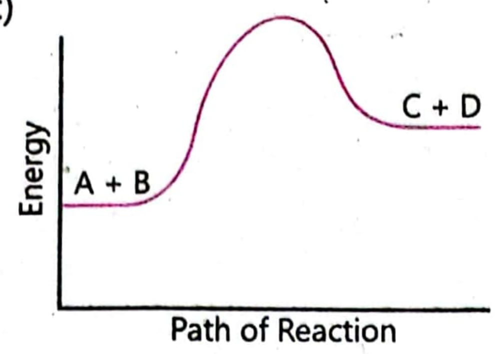
(a)



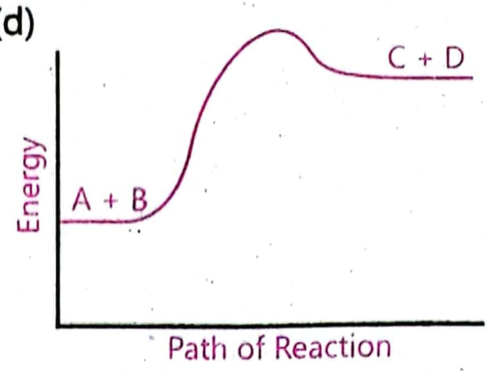
(b)



(c)



(d)



(vi) What does it show when a chemical reaction is exothermic?

(a) It shows the bonds which break are weaker than those are formed.

(b) It shows the bond which break are stronger than those are formed.

(c) Exothermic nature of the reaction is not concerned with bond formation or bond breakage.

(d) It shows that the reactants are more stable than the products.

(vii) When NaOH and HCl are mixed the temperature increases. The reaction is:

(a) endothermic with a positive enthalpy change.

(b) endothermic with a negative enthalpy change.

(c) exothermic with a positive enthalpy change.

(d) exothermic with a negative enthalpy change.

(viii) The average bond dissociation energy for the C-H bond is 412 kJ mol^{-1}. Which of the following process will have enthalpy change close to 412 kJ mol^{-1}?

(a) $CH_{4(g)} \longrightarrow C_{(g)} + 2H_{2(g)}$

(b) $CH_{4(g)} \longrightarrow C_{(g)} + 2H_{2(g)}$

(c) $CH_{4(g)} \longrightarrow C_{(g)} + 4H_{(g)}$

(d) $CH_{4(g)} \longrightarrow CH_{3(g)} + H_{(g)}$

(ix) The average bond energies for O - O and O = O are 146 and 496 kJ mol^{-1} respectively. Find the enthalpy in kJ for the following reaction?

$$H-O-O-H \longrightarrow H-O-H_{(g)} + \frac{1}{2} O\;O_{(g)}$$

(a) -102kJ

(b) +102kJ

(c) +350kJ

(d) +394kJ.

(x) Why does the following exothermic reaction not occur?

$$C \text{ (Diamond)} \longrightarrow C \text{ (Graphite)} \qquad \Delta H = -3 kJ mol^{-1}$$

(a) Structure of diamond is more stable than that of graphite.

(b) Diamond has strong covalent bonds than does the graphite.

(c) The change from diamond to graphite has high activation energy.

(d) Density of graphite is less then that of diamond.

2. Questions for Short Answers

i. What is the difference between enthalpy and enthalpy change?

ii. Why is breaking of a bond an endothermic process?

iii. Depict the transition state for the following reaction.

$$H_2 + Cl_2 \longrightarrow 2HCl$$

iv. Draw the reaction profiles for two exothermic reactions one of which moves faster than the other.

v. What is the role of glycogen in our body?

3. Constructed Response Questions

i. Physical changes which usually occur around us are given in the table. Write down whether they are exothermic or endothermic.

Physical change	Exothermic or Endothermic	Physical change	Exothermic or Endothermic
Conversion of hydrated salt into anhydrous salt		Conduction of electricity by metals	
Burning of paper		Dissolving ammonium chloride in water	
Vapourizing liquid nitrogen		Formation of rain from clouds	
Evaporation of dry ice		Dissolving sodium carbonate in water	

ii. Explain why the reaction between atmospheric gases oxygen and nitrogen does not take place under normal conditions? But in the presence of lightning these gases react to give NO. The reaction stops as soon as the lightning stops.

iii. A reaction between natural gas (CH_4) and atmospheric oxygen does not take place when you mix them. As soon as you show a burning match stick, the reaction starts immediately and then it continues until one or both of the reactants is/are used up. Explain.

4. Descriptive Questions

i. Find out the enthalphy change of the following reaction using the given data.

$$N_2 + O_2 \longrightarrow 2NO$$

Bond dissociation energy of N_2 = 958.38 kJ/mol

Bond dissociation energy of O_2 = 498 kJ/mol

Bond formation energy of NO = -626 kJ/mol

ii. Explain the difference between the terms heat and enthalpy.

iii. Explain why formation of a bond is always an exothermic process.

iv. Explain the role of lipids in our body.

v. Explain the following terms.

Activation energy, Transition state, Aerobic respiration

5. Investigative Questions

i. Why is it essential to cook some of the food items while others we can eat without cooking?

ii. Why does fireworks look spectacular? What type of chemical compounds undergo chemical reactions during this activity?

Chapter 6 Equilibria

Student Learning Outcomes

After studying this chapter, students will be able to:

- Recognize that reversible reactions are shown by symbol and may not go to completion
- Describe how changing the physical conditions of a chemical equilibrium system can redirect reversible reactions (Some examples can include: **a.** effect of heat on hydrated compounds **b.** addition of water to anhydrous substances in particular copper (II) sulfate and cobalt (II) chloride
- State that reversible reactions can achieve equilibrium in a closed system when rate of forward and backward reactions are equal.

6.1 Reversible and Irreversible Changes

In a chemical reaction, the reactants react to give the products. The reaction will continue until all the reactants or one of the reactants is converted into product. For example, the following reaction takes place immediately in an aqueous solution to give the products.

$$NaCl_{(aq)} + AgNO_{3(aq)} \longrightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

The reaction goes to completion and if stoichiometric amounts of the reactants are used then no reactant species are present at the end of the reaction. Such a reaction is called an **irreversible reaction.** It moves in the forward direction only.

In the majority of chemical reactions, however, the reaction does not go to completion. The products of the reaction react among themselves to give back the reactants under the same conditions.. Such a reaction moves in both the forward and backward directions under the same conditions. The reactants react to give the products and the products, in turn, react to give back the reactants. The reaction is called a **reversible reaction** and it is denoted by a symbol $\rightleftharpoons$ which has two half arrowheads one pointing in each direction. For example,

$$N_{2(g)} + 3H_{2(g)} \underset{Fe}{\overset{400°C,\ 200\ atm}{\rightleftharpoons}} 2NH_{3(g)}$$

In this reaction, one mole of nitrogen gas reacts with three moles of hydrogen gas under the given reaction conditions in a closed container to give two moles of ammonia gas. After its formation, the ammonia gas decomposes to give the reactants back. The reaction never goes to completion. At any time, all the three species are simultaneously present in the reaction mixture.

A reversible reaction, however, goes to completion if either one of the products is withdrawn from the reaction mixture or being a gas, it escapes into the atmosphere. For example, calcium carbonate is decomposed by heating at a particular temperature.

$$CaCO_{3(s)} \xrightleftharpoons{\text{Heat}} CaO_{(s)} + CO_{2(g)}$$

If the above reaction is carried out in an open container, the carbon dioxide gas will escape into the atmosphere as soon as it is formed and the reaction is forced to go to completion. If, on the other hand, the reaction is performed in a closed container, the carbon dioxide will react with calcium oxide to give back the reactants.

Like chemical changes, physical changes may also be reversible in nature. Copper sulphate pentahydrate ($CuSO_4$ $5H_2O$) is an important salt which is blue in colour. When this salt is heated strongly, its colour changes to white. This physical change involves the following equilibrium.

$$\underset{\text{Hydrated Copper sulphate}}{CuSO_4 \cdot 5H_2O_{(s)}} \xrightleftharpoons{\text{Heat}} \underset{\text{Anhydrous Copper sulphate}}{CuSO_{4(s)}} + 5H_2O_{(g)}$$

When white anhydrous copper sulphate absorbs moisture from the atmosphere, it will turn blue again.

Similar to this, when cobalt chloride hexahydrate ($CoCl_2$ $6H_2O$), which is pink in colour, is heated it is converted to anhydrous $CoCl_2$ which is blue in colour.

$$\underset{\text{Hydrated Cobalt chloride (Pink)}}{CoCl_2 \cdot 6H_2O_{(s)}} \xrightleftharpoons{\text{Heat}} \underset{\text{Anhydrous Cobalt chloride (Blue)}}{CoCl_{2(s)}} + 6H_2O_{(g)}$$

In the reverse reaction, the anhydrous cobalt chloride absorbs less moisture, it is first converted into a dihydrate which is purple in colour. This dihydrate then further absorbs four more water molecules to become a hexahydrate which is pink in colour.

$$CoCl_2 \underset{}{\overset{\text{Limited moisture}}{\rightleftharpoons}} CoCl_2 \cdot 2H_2O \overset{\text{Moisture}}{\rightleftharpoons} CoCl_2 \cdot 6H_2O$$

Anhydrous Cobalt chloride (Blue) — Cobalt chloride dihydrate (Purple) — Cobalt hexahydrate (Pink)

Activity

Take a few grams of blue coloured copper sulphate in a dry test tube. First heat it gently and then strongly for some time.

Note down the observation. Let the test tube cool down. Again note down the observation.

6.2 Dynamic Equilibrium

If a reversible reaction is started by mixing the reactants, the reaction moves in the forward direction only. After some time when enough concentrations of the products are built up, they react to give back the reactants in the reverse reaction.

The reaction will keep on going in both the directions until the rate of forward reaction becomes equal to the rate of backward reaction. In other words, the number of reactant molecules which will disappear as a result of forward reaction becomes equal to the number of reactant molecules which will form as a result of the **reverse** reaction. The same will be true for the product molecules. At this stage, the reaction is said to be in a state of **chemical equilibrium.** It appears as if nothing is going on in the reaction vessel as the concentrations of both reactant and product molecules do not undergo any change at this stage. Since the reaction did not cease at this **state of equilibrium**, rather it keeps on going in both the directions, this state is called **dynamic equilibrium.** The concentrations of reacting species (reactants and products) remain constant at equilibrium.

Interesting Information!

Vast deposits of coal are available in Thar, Sindh. This coal can be used to generate electricity. When coal is made to react with steam, CO and H_2 are produced. These products then react by a reversible reaction called catalytic methanation to yield methane.

$$C_{(s)} + H_2O_{(g)} \longrightarrow \underbrace{CO_{(g)} + H_{2(g)}}_{\text{Water gas}}$$

$$CO_{(g)} + 3H_{2(g)} \overset{\text{Catalyst}}{\rightleftharpoons} CH_{4(g)} + H_2O_{(g)}$$

Exercise

1. Elaborate an example of dynamic equilibrium which exists in this world between the three physical states of water.
2. Dinitrogen tetraoxide (N_2O_4) is a colourless gas. It slowly changes to brown coloured nitrogen dioxide (NO_2) at 100° C. Predict how the colour of the mixture will change if N_2O_4 is kept in a sealed flask at 100°C.

$$\underset{\textbf{Colourless}}{N_2O_4} \xrightleftharpoons{100^\circ C} \underset{\textbf{Brown}}{2NO_{2(g)}}$$

The time a reaction will take to attain the state of dynamic chemical equilibrium depends upon the nature of the reaction and the conditions at which the given reversible reaction is performed.

For a dynamic equilibrium to be set up, the rate of the forward reaction must be equal to the rate of backward reaction. This does not happen instantly and for very slow reaction, it may take years. Both the rates of formation and the decomposition of ammonia are reasonably fast at around 400°C in the presence of a catalyst. This reaction will reach the equilibrium state within minutes after the start of reaction.

The following equilibrium reaction takes 4 – 5 seconds to reach at the point of equilibrium.

$$2H_2O_{(l)} \xrightleftharpoons{\text{Electric current}} 2H_{2(g)} + O_{2(g)}$$

6.3 Changing the Physical Conditions of a Chemical Reaction

If a given reversible reaction has attained the state of dynamic equilibrium, it will remain in this state for infinite time unless it is somehow disturbed. A reversible chemical system may be disturbed in the following possible ways.

1. Adding or withdrawing one or more of the reacting species
2. Adding or withdrawing one or more of the product species
3. Changing the temperature of the reaction
4. Effect of the presence of a catalyst on a reversible reaction
5. Changing the pressure of the reaction if it involves reactants or the products in the gaseous state

Consider the following reversible reaction at equilibrium.

$$N_{2(g)} + 3H_{2(g)} \xrightleftharpoons[\textbf{Catalyst}]{\textbf{400°C}} 2NH_{3(g)}$$

The concentrations of all the participating chemicals will be constant at the state of equilibrium. At this stage if we add more N_2 gas in the mixture, its concentration will increase and the reaction will no longer maintain its state of equilibrium. To restore the equilibrium state again, nitrogen will react with hydrogen to produce more ammonia. This change will go on until a new state of equilibrium is reached at which the concentrations of all the species will again become constant. These new concentrations will, however, be different from the concentrations of the earlier equilibrium state.

Now let us disturb the equilibrium again by withdrawing some of the ammonia gas formed. As a result, its concentration will decrease. To restore the equilibrium state, more nitrogen and hydrogen will react to produce ammonia. When the state of equilibrium is reached again, the concentrations of all the species shall again become constant.

Effect of Changing the Temperature on the State of Equilibrium

The formation of ammonia is exothermic in the forward direction and hence this reaction will be endothermic in the reverse direction.

$$N_{2(g)} + 3H_{2(g)} \xrightleftharpoons{\text{Heat}} 2\,NH_{3(g)} \qquad \Delta H = -92.4\ \text{kJ/mole}$$

If this reaction is at equilibrium and its temperature is increased, the state of the equilibrium will be disturbed again. The ΔH of this reaction is negative. This means that the total energy of the system containing N_2 and H_2 is higher than that of ammonia. The increase in temperature of this reaction at equilibrium will push the reaction in the backward direction i.e. the reactants side. Decreasing the temperature will drive the equilibrium to the forward direction.

Effect of Change of Pressure on the Reaction at Equilibrium

Change of pressure will disturb the equilibrium state of only those gaseous reactions in which the number of moles of the reacting gases will be different from the number of moles of the gases being produced. Since the formation of ammonia gas meets such a condition, the state of its equilibrium will be disturbed by changing the pressure exerted on the reaction mixture.

$$\underbrace{\underset{\text{1 mole}}{N_{2(g)}} + \underset{\text{3 moles}}{3H_{2(g)}}}_{\text{4 moles}} \rightleftharpoons \underbrace{\underset{\text{2 moles}}{2NH_{3(g)}}}_{\text{2 moles}}$$

In this reaction, 4 moles of reacting gases are producing two moles of product gas. If this reaction is at equilibrium and the pressure is increased, the equilibrium will be disturbed. To restore this, the reaction will move to that side in which the number of moles are less i.e. forward direction. The formation of ammonia gas is thus favoured at high pressure.

Effect of the Presence of a Catalyst on the Reversible Reaction

At the stage of dynamic equilibrium, the rates of both forward and backward reactions are equal.

A catalyst increases both the rates of forward and back reactions of a reversible reaction. So if a reversible reaction is carried out in the presence of a catalyst it will decrease the time taken by the reaction to attain the state of equilibrium.

Exercise: Let us consider another example of a reversible reaction at equilibrium. Phosphorous pentachloride decomposes according to the following equation.

$$PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$$

According to the equation one mole of the gaseous reactant is giving two moles of the product gases. The reaction is an endothermic reaction. Keeping in view the above description of the reaction, answer the following questions.

i. What will happen if the gas mixture is compressed?
ii. What will happen if we add Cl_2 gas to the equilibrium mixture?
iii. What will happen if the temperature of the reaction is increased?

Interesting Information!

Industrial production of ammonia in Haber Process is a very useful application of the phenomenon of chemical equilibrium. Ammonia gas leads to the formation of an important fertilizer urea. The ability of ammonia gas to be converted into its liquid form easily is used to drive the reaction to completion. In this way, practically 100% conversion of N_2 and H_2 to NH_3 is achieved.

Exercise

The preparation of ethyl acetate is commercially very important because it is used as a thinner in paint industry.

$$\underset{\text{Ethyl alcohol}}{C_2H_5OH_{(l)}} + \underset{\text{Acetic acid}}{CH_3COOH_{(l)}} \xrightleftharpoons{\text{dil. } H_2SO_4} \underset{\text{Ethyl acetate}}{CH_3COOC_2H_{5(l)}} + H_2O_{(l)}$$

One way to get the better yield of the product ethyl acetate is to remove water from the reaction mixture as soon as it is formed. Suggest a suitable method to withdraw water from the reaction mixture.

Key Points

1. Majority of chemical reactions are reversible reactions. Reactants react to give the products and the products, in turn, react or decompose to give back the reactants.
2. A reversible reaction never goes to completion. However, it may be forced to go to completion if one or all the products are withdrawn from the reaction mixture as soon as it is or they are formed.
3. Physical changes may also be reversible in nature.
4. A reversible reaction will keep on going in both the directions until the rate of forward reaction becomes equal to the rate of reverse reaction. At this point the reaction is said to be at a state of chemical equilibrium.
5. Since both the forward and reverse reactions keep on going at the state of chemical equilibrium, it is called a dynamic equilibrium. The concentrations of all the reacting species remain constant at equilibrium.
6. The time a reaction takes to attain the state of dynamic equilibrium depends upon the nature of reaction and the conditions at which the reversible reaction is performed.
7. A reversible chemical system may be disturbed either by adding or withdrawing the reactants or the products and by changing the conditions of temperature, pressure and catalyst.

Exercise

1. Tick (✓) the correct answer.

(i) What will happen if the rates of forward and reverse reactions are very high?
  (a) The equilibrium point will reach very soon.
  (b) The equilibrium point will reach very late.
  (c) The reaction will not attain the state of dynamic equilibrium.
  (d) The reaction will be practically irreversible.

(ii) Predict which components of the atmosphere react in the presence of lightning.
  (a) N_2 and H_2O (b) O_2 and H_2O
  (c) CO_2 and O_2 (d) N_2 and O_2

(iii) An inorganic chemist places one mole of PCl_5 in a container A and one mole of each Cl_2 and PCl_3 in container B. Both the containers were sealed and heated to the same temperature to reach the state of equilibrium. Guess about the composition of mixtures in both the containers.
  (a) Both the containers will have the same composition of mixtures.
  (b) Container A will have more concentration of PCl_3 than B.
  (c) Container A will have less concentration of PCl_3 than B.
  (d) Both the containers will have zero concentration of its reactants.

(iv) CaO or lime is used extensively in steel, glass and paper industries. It is produced in an exothermic reversible reaction by the decomposition of lime ($CaCO_3$). Choose the conditions to produce maximum amount of lime.
  (a) Heating at high temperature in a closed vessel
  (b) Heating at high temperature in an open vessel
  (c) Cooling it in a closed vessel
  (d) Cooling it in an open vessel

(v) What condition should be met for the reversible reaction to achieve the state of equilibrium?
  (a) All the reactants should be converted into the products.
  (b) 50% of the reactants should be converted into products.
  (c) The concentrations of all the reactants and the products should become constant.
  (d) One of the products should be removed from the reaction mixture.

(vi) Why the gas starts coming out when you open a can of fizzy drink?

(a) Because the solubility of the gas increases

(b) Because the gas is insoluble in water

(c) Because the gas is dissolved under pressure hence it comes out when pressure is decreased

(d) Because the solubility of the gas decreases at high pressure

(vii) The following reaction is performed in a closed vessel.

$$CaCO_{3\,(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$$

How the equilibrium will be affected if you increase the pressure?

(a) The forward reaction will be favoured

(b) The backward reaction will be favoured

(c) No effect on backward reaction

(d) No effect on forward or backward reaction

(viii) When a reaction will become a reversible one?

(a) If the activation energy of the forward reaction is comparable to that of backward reaction

(b) If the activation energy of the forward reaction is higher than that of backward reaction

(c) If the activation energy of the forward reaction is lower than that of backward reaction

(d) If the enthalpy change of both the reactions is zero.

(ix) Is reversible reaction useful for preparing compounds on large scale?

(a) No (b) Yes

(c) They are useful only when equilibrium lies far to the right side

(d) They are useful only when equilibrium lies far to the left side

(x) What will happen to the concentrations of the products if a reversible reaction at equilibrium is not disturbed?

(a) They will remain constant

(b) They will keep on increasing

(c) They will keep on decreasing

(d) They will remain constant for some time and then start decreasing

2. Questions for Short Answers

i. How is dynamic equilibrium different from the static equilibrium?

ii. How the following reversible reaction will be affected if its temperature is increased?

$$2H_2O_{(l)} \xrightleftharpoons{Electricity} 2H_{2(g)} + O_{2(g)}$$

iii. How can you get the maximum yield in a reversible reaction?
iv. How can you decrease the time to attain the position of equilibrium in a reversible reaction?
v. What is the effect of increasing pressure on the following reaction at equilibrium?

$$N_{2(g)} + O_{2(g)} \rightleftharpoons 2NO_{(g)}$$

3. Constructed Response Questions

I. Why are some reactions irreversible while others are reversible?
ii. Why are combustion reactions generally irreversible?
iii. Can you make an irreversible reaction reversible and vice versa?
iv. How do you know if a reaction is reversible or irreversible?
v. Do the phase changes in water (solid to liquid, liquid to vapour) reversible or irreversible?

4. Descriptive Questions

i. How can you drive the reversible reaction at equilibrium?
(a) in the forward direction (b) in the backward direction
ii. Explain how the rates of forward and backward reactions change when the system approaches equilibrium.
iii. Describe the effect of a catalyst on the reversible reaction.
iv. How can a reversible reaction be forced to go to completion?
v. How does change in temperature affect the reaction at equilibrium?

5. Investigate

I. Study the effect of heat on hydrated copper sulphate Why does this salt look coloured and why does it lose colour upon heating?
ii. Synthesis of ammonia gas is very important industrially because it is used in the preparation of urea fertilizer. Explain the conditions you will use to get the maximum yield of ammonia.

Chapter 7

Acid Base Chemistry

Student Learning Outcomes

After studying this chapter, students will be able to:

- Define Bronsted-Lowry acids as proton donors and Bronsted -Lowry bases as proton acceptors
- Recognize that aqueous solutions of acids contain H^+ ions and aqueous solutions of alkalis contain OH^- ions
- Define a strong acid and base as an acid or base that completely dissociates in aqueous solution and weak acid and base that partially dissociates in aqueous solution. (Some examples include: Student writing symbol equations to show these for hydrochloric acid, sulphuric acid, nitric acid, and ethanoic acid.)
- Formulate dissociation equations for an acid or base in aqueous solution.
- Recognize that bases are oxides or hydroxides of metals and that alkalis are water-soluble bases
- Describe the characteristic properties of acids in terms of their reactions with metals, bases and carbonates
- Identify the characteristic properties of bases in terms of their reactions with acids and ammonium salts
- Define acid rain.
- Discuss effects of acid rain and relate them with properties of acids.

7.1 Acids and Bases

Acids and bases have been known to mankind since centuries. Acids have been known for their sour taste like the taste of lemon and ability to change the colour of litmus paper from blue to red. There are many substances which contain acids and hence taste sour, such as curd, tamarind, lemon and lime, etc.

Common examples of acids include acetic acid, hydrochloric acid, nitric acid, sulphuric acid and tartaric acid. Most of these acids, we come across in everyday life, are available in the form of aqueous solutions as they can be easily dissolved in water.

Acids are divided into two types on the basis of their occurrence: natural acids and mineral acids. Acids which are obtained from natural sources are called natural or organic acids. Mineral or man-made acids are prepared from minerals like sodium chloride or sodium nitrate. Common examples of mineral acids are hydrochloric acid, sulphuric acid and nitric acid.

Some common organic acids and their natural sources are given in Table (7.1)

Table (7.1) Some Common Organic Acids and their Sources

Organic acid	Natural Source
Acetic acid	Vinegar
Ascorbic acid	Amla, Guava
Citric acid	Lemon, Orange
Lactic acid	Sour milk, Curd
Formic acid	Ant sting
Oxalic acid	Tomato
Tartaric acid	Tamarind

In comparison to acids, bases are known for their bitter taste, slippery touch and ability to change the colour of litmus paper from red to blue. An alkali is a base that dissolves in water. Some common examples of alkalis are sodium hydroxide (caustic soda), potassium hydroxide (caustic potash) calcium hydroxide (lime water) and aqueous ammonia (NH_4OH). Both sodium hydroxide and potassium hydroxide are extremely corrosive and can burn your skin.

Metal oxides are also basic in nature because they react with acids to form salts and water. Na_2O is basic oxide because it contains oxide ion, O^{2-}, which is a very strong base with a strong tendency to react with water to produce hydroxide ions.

$$Na_2O_{(s)} + H_2O_{(l)} \longrightarrow 2NaOH_{(aq)}$$

Other examples of basic oxides are calcium oxide (CaO), zinc oxide (ZnO) and magnesium oxide (MgO).

Both acids and bases are known to cancel the properties of each other when mixed together in equal amounts. The reaction is called neutralization reaction. A salt and water are formed as a result of this reaction.

Exercise Name some fruits which contain citric acid.

Interesting Information!

Oxalic acid ($H_2C_2O_4$) is the simplest organic diprotic acid. Its commercial uses include bleaching straw and leather and removing rust and ink stains from fabrics.

Activity 7.1

The following compounds are provided in the form of liquid or in the form of their aqueous solutions. Use blue and red litmus paper strips to show whether these substances are acidic or basic in their aqueous solutions.

Substance	Acidic	Basic
Tap water		
Battery water		
Rain water		
Soap solution		
Tooth paste		
Shampoo		
Bleach		

Exercise

In what ways are mineral acids useful for us?

7.2 Different Concepts of Acids and Bases

Arrhenius Acids and Bases

Although the earlier definitions of acids and bases describe some distinctive features of these substances, yet new and broader definitions were needed to explain their chemical behavior on the molecular level.

Svante Arrhenius, a Swedish Chemist, suggested that acids and bases may be classified in terms of their behavior in water. According to him;

An acid is that substance which dissociates in water to give proton (H^+) or hydroxonium ion (H_3O^+). Some typical Arrhenius acids are HCl, HNO_3, H_2SO_4 and HCN.

$$HCl_{(aq)} \xrightarrow{H_2O} H_3O^+_{(aq)} + Cl^-_{(aq)}$$

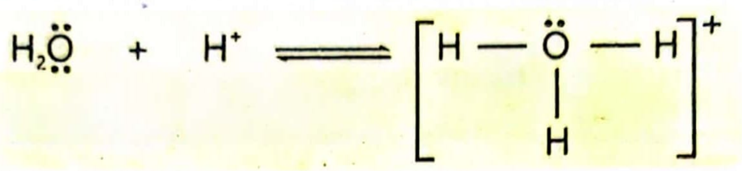
Svante Arrhenius (1859-1927)

Similarly, a base is that substance which dissociates in water to give hydroxyl ions (OH^-).

Some typical Arrhenius bases are NaOH, KOH and $Ba(OH)_2$.

$$NaOH_{(aq)} \xrightleftharpoons{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$$

Water has an essential role to play in Arrhenius concept of acids and bases. Whenever an acid or a base dissociates in water, its molecules participate in reaction by surrounding the resultant proton (H^+) and hydroxyl ion(OH^-). Since proton is very small in size and its charge density is very high, it forms a strong bond with the lone pair of water molecule to give hydroxonium ion, H_3O^+.

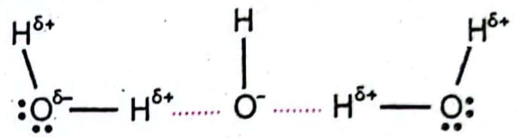


According to some reports upto four water molecules may surround the proton. Similarly, hydroxyl ions are also surrounded by water molecules as shown in the Figure (7.1).

Arrhenius also explained the process of neutralization. According to him when a strong acid and a strong base are dissolved in water, they completely dissociate into ions.

$$HCl_{(aq)} \xrightleftharpoons{H_2O} H_3O^+_{(aq)} + Cl^-_{(aq)}$$

$$NaOH_{(aq)} \xrightleftharpoons{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$$



Arrangement of water molecules around H^+

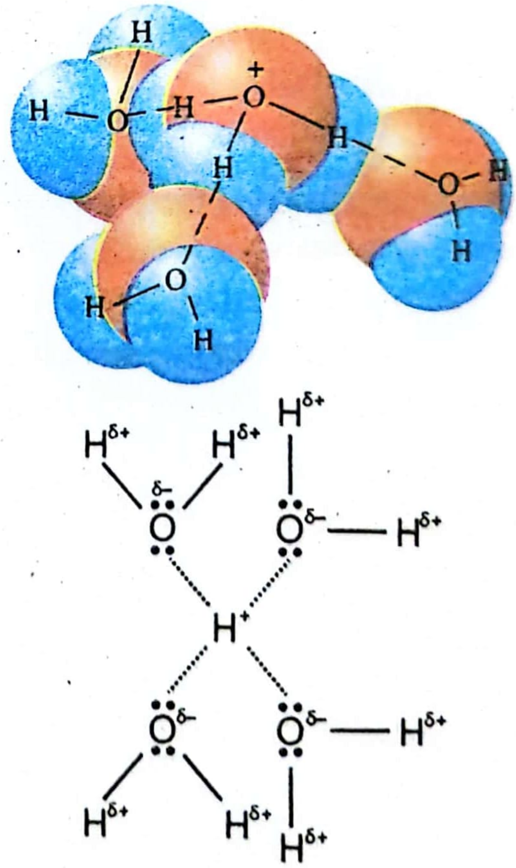


Fig (7.1)

Exercise

How do chloride ions exist in water?

The hydroxonium ion and the hydroxyl ion then react to form water with the evolution of heat.

$$H_3O^+_{(aq)} + OH^-_{(aq)} \longrightarrow 2H_2O_{(l)} + \text{Heat}$$

The salt NaCl that is formed with water does not exist as solid crystals. It remains present in the solution in the form of hydrated spectator sodium ions (Na^+) and chloride ions (Cl^-).

Mineral acids are generally very strong acids. The strength of an acid depends upon the extent to which it is ionized in water. Hydrochloric acid ionizes in water completely giving a large amount of hydroxonium ions in water.

$$HCl_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$$

Sulphuric acid being a diprotic acid ionizes in two steps.

$$H_2SO_{4(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HSO_4^-{}_{(aq)}$$

$$HSO^-_{4(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Nitric acid ionizes in single step.

$$HNO_{3(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$$

Contrary to mineral acids, organic acids ionize upto a very limited extent and hence they are weak acids.

Glacial acetic acid has a percent ionization of only 0.132%.

$$CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$

It means that out of 1000 molecules of acetic acid, only 1.32 molecules dissociate and the rest remain undissociated.

Similarly, the percent ionization of formic acid having concentration 1.50M is 1.06%. Which means 987 molecules out of 1000 remain undissociated.

$$HCOOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HCOO^-_{(aq)}$$

NaOH and KOH are the examples of strong bases because they also ionize completely in water.

$$NaOH_{(aq)} \xrightleftharpoons{H_2O} Na^+ + OH^-_{(aq)}$$

NH_4OH and $Al(OH)_3$ are the examples of weak bases because they only partially ionize in water.

$$NH_4OH_{(aq)} \xrightleftharpoons{H_2O} NH^+_{4\,(aq)} + OH^-_{(aq)}$$

Exercise

Why does ammonium hydroxide only partially ionize in water?

Interesting Information!

Stomach Acidity

Stomach acidity or hyperacidity conditions are a common problem. Most often the problem arises when a person takes fatty and spicy foods which cause more acid to produce in the stomach than required.

Our stomach produces hydrochloric acid to digest the food that we eat. Whenever we eat, cells within the lining of the stomach produce acid. Problem occurs when these cells produce more acid than your stomach needs. When it happens, the person suffers from stomach acidity. The common indication of such a condition is the feeling of burning sensation right below our breast bone. A person may also feel sour taste in mouth and heart burn or pain near the heart area. The uneasy condition may easily be cured by taking weak bases like calcium hydroxide and magnesium hydroxide commonly known as antacids. These antacids remove minor stomach disorders by neutralizing the stomach acid, but the concentration of hydroxyl ions in them is too low to harm the throat or stomach.

7.3 Bronsted – Lowry concepts of Acids and Bases

Arrhenius concept of acids and bases is very simple and easily understandable but it has its own limitations. For example, Na_2CO_3, K_2CO_3 and NH_3 do not contain any hydroxyl group which will get ionized by water but all these compounds behave as bases and yield OH^- in water. Arrhenius definitions also require that water must be present as a solvent in order for a compound to behave as an acid or a base. It was realized that a broader definition for acids and bases was needed to cover all the aspects of the concept.

To remove the limitations of Arrhenius concept, Bronsted and Lowry gave the following definitions of acids and bases:

An acid is a substance that donates a proton.

This definition requires that to behave as an acid a compound must have a proton to donate. The condition of the presence of water during this donation was, however, eliminated. All Arrhenius acids are, thus, Bronsted-Lowry acids as well for example, HCl. It dissociates in water to give H^+ and Cl^-. It also donates H^+ to H_2O forming H_3O^+.

Johannes N. Bronsted (1879-1947) Thomas M. Lowry (1874-1936)

A base is a substance that accepts a proton.

For example, OH^-, CN^-, NH_3 and Cl^- are all bases because they have the ability to accept a proton. Note that except OH^- all other species are not Arrhenius bases. All Arrhenius bases are, however, Bronsted-Lowry bases as well.

All Bronsted-Lowry acids and bases are not Arrhenius acids and bases. NH_4^+ is not Arrhenius acid and NH_3 is not Arrhenius base.

According to Bronsted-Lowry, an acid base reaction is that reaction in which a proton is transferred from a proton donor to its acceptor. This reaction may take place in gas phase or in the presence of any solvent.

Consider the following reaction between hydrogen chloride gas and liquid water.

$$\underset{\text{Acid}}{HCl_{(g)}} + \underset{\text{Base}}{H_2O_{(l)}} \rightleftharpoons H_3O^+_{(aq)} + Cl^-$$

In this reaction, HCl gas acts as an acid because it donates its proton to water which acts as a base.

Similarly when ammonia gas dissolves in water, a proton is transferred from water to ammonia and ammonium ion is formed.

$$\underset{\text{Base}}{NH_{3(g)}} + \underset{\text{Acid}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Conjugate base}}{OH^-_{(aq)}} + \underset{\text{Conjugate acid}}{NH_4^+{}_{(aq)}}$$

Ammonia is a base while water is an acid in this reaction. Water has the ability to act both as an acid or a base depending upon the other compound with which it reacts.

Water is therefore called an **amphoteric** compound which means a compound that can behave both as an acid and a base.

In the reverse reaction, OH^- is a base because it accepts a proton donated by the acid NH_4^+. In order to differentiate, OH^- is called the conjugate base while NH_4^+ the conjugate acid.

Some other examples of Bronsted-Lowry acids and bases.

$HCN_{(aq)}$	+	$H_2O_{(l)}$	$\rightleftharpoons$	$H_3O^+_{(aq)}$	+	$CN^-_{(aq)}$
Acid		Base		Conjugate acid		Conjugate base
$HI_{(aq)}$	+	$H_2O_{(l)}$	$\rightleftharpoons$	$H_3O^+_{(aq)}$	+	$I^-_{(aq)}$
Acid		Base		Conjugate acid		Conjugate base

Exercise

Give two examples of Bronsted–Lowry bases which are not bases by Arrhenius definition.

7.4 Properties of Acids and Bases

Properties of acids and bases vary depending upon the substances they react with. For example, acids react with metals to produce hydrogen gas whereas bases do not react with metals.

a. Acids give the following three types of reactions.

1. With alkalis or metal oxides, they form salts and water.

$$2HNO_3(aq) + CaO(s) \longrightarrow Ca(NO_3)_2(aq) + H_2O(\ell)$$
$$H_2SO_4(aq) + 2KOH(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(\ell)$$

2. With reactive metals (Mg, Zn) they form salts and evolve hydrogen gas.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$
$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

The less reactive metals Cu, Ag, Au, and Pt do not evolve hydrogen gas with acids.

Activity 7.2

Take a few granules of zinc in a clean test tube. Add dilute H_2SO_4 in it and heat it gently. Identify the gas evolved in this reaction by taking a burning match stick near the mouth of test tube. Note the observations in your notebook.

3. They decompose metal carbonates and hydrogen cabonates evolving carbon dioxide gas.

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + H_2O\,(\ell) + CO_2\,(g)$$

$$2NaHCO_3\,(s) + H_2SO_4\,(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(\ell) + 2CO_2\,(g)$$

b. Properties of Bases

As already mentioned alkalis are those bases which are soluble in water. Examples of alkalis are NaOH, KOH etc. $Ca(OH)_2$ is sparingly soluble in water while $Cu(OH)_2$ is insoluble.

Apart from reacting with acids in neutralization reactions, alkalies also react with ammonium salts and liberate ammonia gas

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl\,(aq) + H_2O\,(\ell)$$

$$Ca(OH)_2 + H_2SO_4(aq) \longrightarrow CaSO_4\,(aq) + 2H_2O\,(\ell)$$

$$Ca\,(OH)_2\,(aq) + 2NH_4Cl\,(aq) \longrightarrow CaCl_2\,(aq) + 2H_2O\,(\ell) + 2NH_3\,(g)$$

$$NaOH\,(aq) + NH_4NO_3\,(aq) \longrightarrow NaNO_3\,(aq) + H_2O\,(\ell) + NH_3\,(g)$$

Activity 7.3

How to clean a blocked drain?

Blocked drains are one of the most common problems that we face every other day. Different blockages require different chemicals to remove them.

One of the ways to clean the drain is to pour half a cup of sodium bicarbonate solution into the drain followed by half a cup of vineger. Cover the drain and wait for thirty minutes. Pour boiling water down the drain.

Caustic chemical drain cleaners are capable of dissolving grease, hair, food and other common blockages. Pour down the caustic cleaner into your drain. Wait for half an hour and then flush your drain with water.

7.5 Acid Rain and its Effects

When rain water has pH between 4.2 – 4.4, it is called acid rain. Burning of fossil fuels release harmful gases in air. These gases, SO_2 and NO_2 when mixed with moisture present in air form acid droplets. These droplets then fall on the ground as acid rain.

$$2SO_{2(g)} + O_{2(g)} \longrightarrow 2SO_{3(g)}$$
$$SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(g)}$$
$$2NO_{2(g)} + H_2O_{(l)} \longrightarrow HNO_{3(g)} + HNO_{2(g)}$$

Effects of Acid Rain

Acid rain causes a number of adverse effects on soil, plants, aquatic life and human-made structures.

Acid rain makes soil more acidic. It dissolves and washes away nutrients present in the soil which are needed by plants. Many plants cannot live or grow in an acidic soil. It can damage vegetation and plants.

Acid rain can make water of the water bodies too acidic for aquatic animals to live in. Due to this , many lakes, rivers, ponds and streams no longer have fish.

Acid rain and the dry deposition of acidic particles damage buildings, statues, automobiles and other structures made of stone and metal.

Key Points

1. Acids are those compounds which have a sour taste and which turn blue litmus red. They also give hydroxonium ions when dissolved in water.
2. Bases or alkalis are those compounds which are bitter in taste, have a slippery touch and which change red litmus blue. They also form hydroxide ions when dissolved in water.
3. In a neutralization reaction, an aqueous solution of an acid reacts with an aqueous solution of a base to give salt and water.
4. According to Arrhenius definitions, acids give protons in water and bases give hydroxide ions in water.
5. Bronsted – Lowry define acid as a proton donor and base as proton acceptor.
6. Generally acids dissolve metals with the evolution of hydrogen gas. They also decompose carbonates and hydrogen carbonates.
7. Generally bases react with ammonium salts to evolve ammonia gas.

Exercise

1. Tick (✓) the correct answer.

(i) Which acid is not used as a food or mixed with food?

(a) Tartaric acid (b) Ascorbic acid
(c) Citric acid (d) Formic acid

(ii) While baking, which gas is responsible for raising the bread and making it soft?

(a) Oxygen (b) Carbon dioxide
(c) Nitrogen (d) Carbon monoxide

(iii) Predict the main characteristics of the reactions of metals with acids.

(a) Metals are dissolved
(b) Metals are converted into salts
(c) Hydrogen gas is evolved
(d) All the above mentioned characteristics are true

(iv) How many hydroxide ions, calcium hydroxide will release in water?

(a) 1 (b) 2 (c) Zero (d) 3

(v) In a neutralization reaction between KOH and H_3PO_4, how many molecules of KOH will react with one molecule of H_3PO_4?

(a) 2 (b) 1 (c) 3 (d) 4

(vi) Which acid is used in the preparation of soap?

(a) Tartaric acid (b) Citric acid
(c) Stearic acid (d) Oxalic acid

(vii) Which compound is formed when SO_2 is dissolved in water?

(a) SO_3 (b) H_2SO_3 (c) H_2SO_4 (d) $H_2S_2O_7$

(viii) Which of the following contains oxalic acid?

(a) Tomato (b) Orange (c) Tamarind (d) Sour milk

(ix) Which compound in the following reaction is behaving as a conjugate base?

$$CH_3COOH_{(aq)} + H_2O_{(l)} \longrightarrow CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

(a) CH_3COOH (b) H_2O (c) CH_3COO^- (d) H_3O^+

(x) When a chemical reaction is carried out with a substance Z; a gas is produced which turns red litmus paper blue. What is the reaction?

(a) Reaction of an acid with a metal carbonate
(b) Reaction of an acid with metal hydrogen carbonate
(c) Reaction of an alkali with an acid
(d) Reaction of an alkali with ammonium salt

2. Questions for Short Answers

i. Choose Arrhenius Acids among the following compounds.

HF, $\overset{+}{N}H_4$, H_2SO_3, SO_3, H_2S, H_2O

ii. How does calcium metal react with dilute H_2SO_4?

iii. Which salt is formed when HCl reacts with $BaCO_3$?

iv. How will you justify that HSO^-_4 is a Bronsted – Lowry acid?

v. Why is HCl not edible although it is present in the stomach and responsible for digestion of food?

3. Constructed Response Questions

i. What chemical name will you give to soap as a compound?

ii. In the presence of a drop of an acid, water is known to ionize as follows:

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

In your opinion, which name will be suitable for water: an acid, a base or both?

iii. Why does Na_2CO_3 behave like a base in water?

iv. Is $NaHCO_3$ a base or an acid? Justify your answer.

v. What is the difference between a strong acid and a concentrated acid?

4. Descriptive Questions

i. Explain Arrhenius concept of acids and bases.

ii. Compare Arrhenius and Bronsted – Lowry concepts of acids and bases.

iii. How does sulphuric acid react with the following compounds?

NH_4Cl, NH_4NO_3, MgO, $MgCO_3$

iv. What happens when a base reacts with a non-metallic oxide. What do you infer about the nature of non-metallic oxide?

v. State the reason of showing acidic character by both dry HCl gas and HCl solution in water.

vi. Differentiate between an acid and its conjugate base.

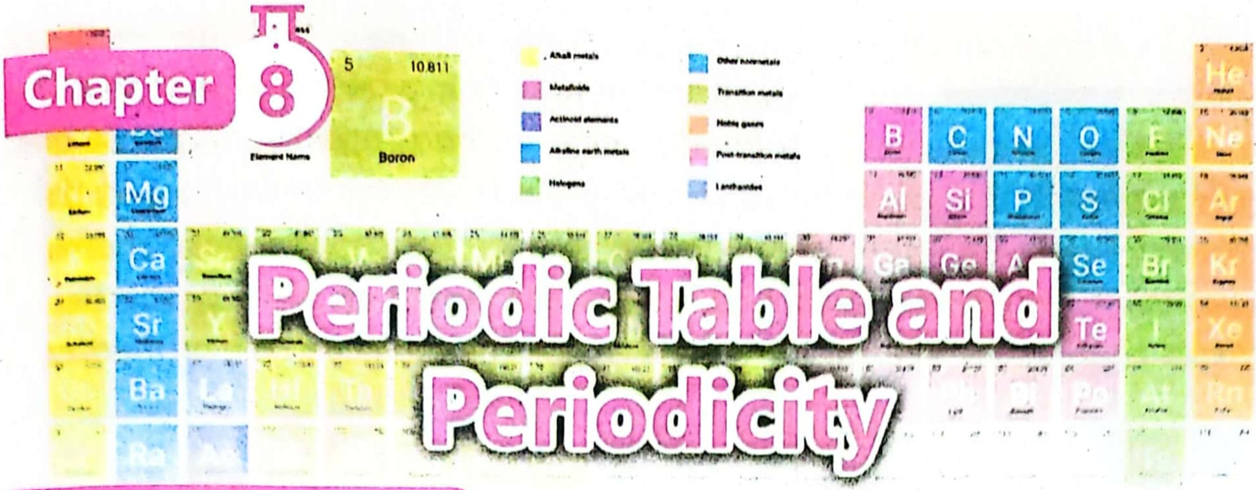
5. Investigative Questions

i. Acids play significant roles within human body. Comment on this statement.

ii. What is observed when CO_2 is passed through lime water (i) for a short duration (ii) for a long duration?

Chapter 8

Periodic Table and Periodicity



Student Learning Outcomes

After studying this chapter, students will be able to:

- Define the periodic table as an arrangement of elements in periods and groups, in order of increasing proton number/atomic number
- Identify the group or period or block of an element using its electronic configuration (only the idea of subshells related to the blocks can be introduced)
- Explain the relationship between group number and the charge of ions formed from elements in the group in terms of their outermost shells
- Explain similarities in the chemical properties of elements in the same group in terms of their electronic configuration
- Identify trends in group and periods, given information about the elements, including trends for atomic radius, electron affinity, electronegativity, ionization energy, metallic character, reactivity and density
- Use terms alkali metals, alkaline earth metals, halogens, noble gases, transition metals, lanthanides and actinides in reference to the periodic table
- Predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity
- Deduce the nature, possible position in the Periodic Table and the identity of unknown elements from given information about their physical and chemical properties

8.1 Modern Periodic Table

The modern periodic table is based upon the arrangement of elements according to increasing atomic number. When the elements are arranged according to ascending order of their atomic numbers from left to right in a horizontal row, properties of elements are found repeating after regular intervals. This results in the form of a table in which elements of similar properties are placed in the same vertical columns.

The horizontal rows of elements in the periodic table are called periods while the vertical columns are called groups.

In the modern periodic table, the electronic configuration of the elements continues changing when we move from left to right in a period. Due to this, the elements in a period show a gradual change in the properties. Against this, elements present in a group possess similar electronic configurations in their outermost shells. Therefore, the elements in a group show similar chemical properties. As a whole the periodic table shows the repetition of the properties of elements after regular intervals. This trend is called periodicity in the properties.

8.2 Salient Features of Modern Periodic Table

Periods

The horizontal rows in the modern periodic table are called periods. There are seven periods in total. Each period except the first starts with an alkali metal and ends at a noble gas. Each period also represents the completion of a shell. Since the number of electrons to be accommodated in a particular shell is fixed, the number of elements in a period is also fixed. The detail about the periods is given in Table (8.1).

Table (8.1) Periods in the Periodic Table

Period No.	Name of the Period	Number of Elements	Number of Shell being filled
1st	Short	2	1st
2nd	Normal	8	2nd
3rd	Normal	8	3rd
4th	Long	18	4th
5th	Long	18	5th
6th	Very long	32	6th
7th	Very long	32	7th

In 6th and 7th periods, two series of fourteen elements each have been accommodated. Because of the space problem, these two series were placed at the bottom of the periodic table to keep it in a manageable and presentable form. The first series starts after lanthanum (La = 57) and it is called **lanthanides**. The second series starts after actinium (Ac = 89) and it is called **actinides**. Together the elements present in these two series are also called **rare earths or f-block elements**.

Groups

The vertical columns present in the periodic table are called groups. There are in total eighteen groups. Elements present in a group resemble one another in their chemical properties since they contain the same number of electrons in their outermost shells. Elements present in a group are also called a family and each group has also been given a family name. The distribution of electrons in the outermost shells (electronic configuration) and other information about the groups are given in the following Table (8.2).

Table (8.2) Electronic Configurations of Elements In The Outermost Shell

Group No.	Family Name	Electronic Configuration In the Outermost Shell
1	Alkali metals	ns^1
2	Alkaline earth metals	ns^2
3 to 12	Transition metals	$nd^x\ ns^2$
13	Boron family	$ns^2\ np^1$
14	Carbon family	$ns^2\ np^2$
15	Nitrogen family	$ns^2\ np^3$
16	Oxygen family	$ns^2\ np^4$
17	Halogen family	$ns^2\ np^5$
18	Noble gases	$ns^2\ np^6$

The groups 1 to 2 and 13 to 17 contain the normal elements. In the normal elements, all the inner shells are completely filled while the outermost shell is incomplete. The groups 3 to 12 are called transition elements and in these elements the inner sub-shells are in the process of completion.

Elements present in the periodic table are also classified into blocks. It depends upon the type of the subshell which is being filled; s, p, d and f. Elements of group 1 and 2 are called **s-block elements** because in them s-subshell of the outermost shell is being filled. Similarly, elements present in groups 13 to 18 are called **p-block elements** because p-subshell is filled in these elements. The **d-block elements** lie between the s and p blocks, while **f-block elements** in the form of two rows lie at the bottom of the periodic table.

Interesting Information!

Mendeleev arranged elements in his periodic table in the ascending order of their atomic masses. Later on, Moseley arranged the elements in ascending order of their atomic numbers.

PERIODIC TABLE OF THE ELEMENTS

30 Zn Zinc: Atomic Number, Element Symbol, Element Name

1 H Hydrogen 1.008																	2 He Helium 4.0026
3 Li Lithium	4 Be Beryllium											5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
11 Na Sodium	12 Mg Magnesium											13 Al Aluminium	14 Si Silicon	15 P Phosphorus	16 S Sulfur	17 Cl Chlorine	18 Ar Argon
19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
55 Cs Caesium	56 Ba Barium	57 La Lanthanum	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Re Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon
87 Fr Francium	88 Ra Radium	89 Ac Actinium	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt	110 Ds	111 Rg	112 Cn Copernicium	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium
90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium

Interesting Information!

E.G. Mazurs collected 700 different published versions of the periodic table. Many forms retain the regular rectangular structure. Some forms had spirals, circles and triangular shapes.

Exercise

(1) The electronic configuration of the outermost shell of an element is s^2p^3. Find out period number and the group number of the element. In which block will you place this element?

(2) What is the group of the element having eight electrons in its outermost shell? In which physical state does this element exist?

(3) An element belongs to 16th group and it is a gas. To which period does it belong?

8.3 Similarities in the Chemical Properties of Elements in the Same Group

The chemical properties of elements depend largely upon the number of electrons present in their outermost shells. Since in a group of the periodic table all the elements have the same number of electrons in the outermost shell, they are expected to show similar chemical properties.

All elements of group I have one electron in their outermost shells, so they show a strong tendency to lose their electron forming cations. They are thus known as electropositive metals. These metals react vigorously with water producing hydrogen and giving alkalies in the solution.

$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow \underset{\text{Alkali}}{2NaOH_{(aq)}} + H_{2(g)}$$

Oxides of these metals are also strongly basic in nature. They are readily soluble in water giving alkalies.

$$K_2O_{(s)} + H_2O_{(l)} \longrightarrow 2KOH_{(aq)}$$

Alkali metals also react with halogens giving halides.

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

The reactivity of alkali metals gradually increases down the group.

The second group elements also show a tendency to lose both of their outermost electrons forming dipositive ions.

$$Mg_{(s)} \longrightarrow Mg^{2+}_{(s)} + 2e^-$$

The tendency to lose electrons down the group increases due to gradual increase in their atomic sizes. The oxides of these metals are also basic in nature and form alkalies in water.

$$CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2\,(aq)}$$

All elements in group 17 have s^2, p^5 configuration in their outer shells. They show a strong tendency to gain one electron to become an anion. They are called electronegative elements. These elements are very reactive non-metals and form salts with metals. Hence they are called halogens (salt forming). Unlike alkali metals, the reactivity of halogens decreases down the group.

$$Cl_{2\,(g)} + 2e^- \longrightarrow 2Cl^{1-}_{(g)}$$

$$Ca_{(s)} + Cl_{2\,(g)} \longrightarrow CaCl_{2\,(s)}$$

Elements present in group 16 have s^2,p^4 configuration in their outermost shells, so they have a tendency to accept two electrons to form a dinegative ion. Oxygen behaves as a strong electronegative element.

$$O_{2\,(g)} + 2e^- \longrightarrow 2O^{2-}_{(g)}$$

Relationship between Group Number and Charge of Ions

When we move from left to right in the periodic table, the main-group elements tend to form cations having a charge equal to the group number. For example, group 1 elements form 1+ ions, group 2 elements forms 2+ ions and group 3 elements form 3+ ions. The number of charges on the cations also correspond to the number of electrons present in their outermost shells.

When we move from right to left in the periodic table, elements often form anions with a negative charge equal to the number of group towards the left side of the noble gases. For example, group 17 elements (which are located one group towards left to the noble gases) form ions with one negative charge. Similarly the group 16 elements form ions with two negative charges. The negative charges present on these ions correspond to number of electrons which these groups need to complete their octets.

8.4 Variation of Periodic Properties in Periods and Groups

Periodicity in the properties of the elements occur due to recurrence of similar electronic configuration in the outermost shells. These properties include:

i) Atomic radius
ii) Ionization Energy
iii) Electron affinity
iv) Electronegativity

Atomic Radius

The electron cloud of an atom has no definite limit. Because of this, the size of an atom cannot be defined exactly. However, it is possible to measure the radius of an atom when it is bonded to an identical atom.

Atomic radius is defined as half the distance between the nuclei of the two identical bonded atoms. It is expressed in pm($1pm = 10^{-12}m$). For example, the distance between the nuclei of two bonded carbon atoms is 154 pm. Half of this distance i.e. 77pm is therefore the radius of carbon atom. This is also called covalent radius of carbon atom.

Variation of Atomic Radius in Periods

When we move from left to right in a period, the size of atoms decreases generally. It is because as we go from lithium (Li) to neon (Ne) in the second period, we are adding electrons to the outermost shell. The charge on the nucleus also increases from +3 to +10. This tends to pull the electrons closer to the nucleus and hence the sizes of atoms decrease from lithium to neon Table (8.4).

Table(8.4) Atomic Radii of Second Period Elements

2nd period elements	Li	Be	B	C	N	O	F	Ne
Atomic Radii (pm)	152	113	88	77	75	73	71	69

Variation of Atomic Radius in Groups

The atomic radii of atoms increase from top to bottom in a group. It is because a new shell is being added in the successive period down the group and the inner electrons shield the outermost electrons from the nucleus Table (8.5).

Table (8.5) Atomic Radii of First Group Elements

First Group Elements	No. of electrons in the Inner Shells	Atomic Radius (pm)
Li	2	152
Na	10	186
K	18	227
Rb	36	248
Cs	54	265

Interesting Information!

Although you might expect atoms to become larger with the increase in their atomic numbers, this does not always occur because the size of an atom is determined by the diameter of its electron shells.

Ionization Energy

Ionization energy is the amount of energy required to remove the most loosely bound electron from the valence shell of an isolated gaseous atom. When there is only one electron present in the valence shell, the energy required to remove it will be called first ionization energy. It is expressed in kJ mol^{-1}. For example, the first ionization energy of sodium atom is +496 kJ mol^{-1}.

$$Na \longrightarrow Na^{+} + e^{-} \qquad \Delta H = +496 \text{ kJ mol}^{-1}$$

When there are more than one electrons in the valence shell, they can be removed one by one providing more and more energy. For example, Mg has two electrons in its outermost shell. It is easier to remove the first electron from magnesium than the second one.

$$Mg_{(g)} \longrightarrow Mg^{+}_{(g)} + e^{-} \qquad \Delta H = +737 \text{ kJ mol}^{-1}$$

$$Mg^{+}_{(g)} \longrightarrow Mg^{2+}_{(g)} + e^{-} \qquad \Delta H = +1450 \text{ kJ mol}^{-1}$$

Ionization energy value is related to the atomic size. The smaller the radius of an atom, the stronger the attraction between the nucleus and the outer electrons and higher the value for ionization energy. The ionization energy values thus increase from left to right in a period and decrease from top to bottom in a group. Table (8.6) and Table (8.7).

Table (8.6) Ionization Energies of Elements of Second Period

2nd period elements	Li	Be	B	C	N	O	F	Ne
Ionization energy (kJ mol^{-1})	520	899	801	1086	1402	1314	1681	2081

Table (8.7) Ionization Energies of Elements of First Group

First Group Elements	Ionization Energy (kJ mol^{-1})
Li	520
Na	496
K	419
Rb	403
Cs	377

Electron Affinity

Electron affinity is the amount of energy change when an electron is added up in the outermost shell of an isolated gaseous atom. For example, 328 $kJmol^{-1}$ energy is released when an electron enters in the fluorine atom.

$$F_{(g)} + e^{-} \longrightarrow F^{1-}_{(g)} \quad \Delta H = -328\ kJmol^{-1}$$

Electron affinity values are also related to the sizes of the atoms. The smaller the size of an atom, the higher the force of attraction with which the nucleus will attract the entering electron and hence higher is the value of electron affinity. Table (8.8)

Table (8.8) Electron Affinities of Group 17 Elements

Group 17 Elements	Electron Affinity ($kJ\ mol^{-1}$)
F	-328
Cl	-349
Br	-324.6
I	-295.2

In a group, the electron affinity values decrease from top to bottom because the sizes of atoms increase down the group. In a bigger atom, the nucleus will attract the incoming electron with a weaker force and hence the electron affinity will also be low.

Things to Know

While first electron affinities can be negative or positive, second electron affinities are always positive.

Electronegativity

Electronegativity of an atom is its electron-attracting ability. It is defined as the force with which an atom attracts the shared pair of electrons towards itself in a bond. Electronegative atoms are those whose outer electrons are attracted by a large nuclear charge. It increases from left to right in a period and decreases from top to bottom in a group. Thus the most electronegative atoms are found at the top right-hand corner of the periodic table. The most electronegative atoms are F,O,N and Cl. Table (8.9) and Table (8.10). The most electropositive elements lie at the bottom left of the periodic table.

Table (8.9) Electronegativities of Elements of Second Period measured on Pauling scale

2nd period elements	Li	Be	B	C	N	O	F
Electronegativity	1.0	1.6	2.0	2.6	3.0	3.4	4.0

Table (8.10) Electronegativities of Elements of Group 17 measured on Pauling scale

17th Group Elements	Electronegativity
F	4.0
Cl	3.2
Br	3.0
I	2.7

Interesting Information!

Electronegativity is one of the most well-known property for explaining why chemical reactions take place.

8.5 Metallic Character and Reactivity

The metallic character is the tendency of an element to lose electrons and form positive ions or cations.

Since the ionization energy decreases down the group, the elements have increased ability to lose electrons. For this reason both the metallic character and reactivity increase down the group.

As we move from left to right in a period, the nuclear charge increases due to a gradual increase in the number of protons in the nucleus. Owing to this the valence electrons are pulled strongly by the nucleus making it difficult for the atoms to lose electrons. Hence the metallic character decreases in a period from left to right.

The chemical reactivity gradually decreases as we move from left to right in a period. For example, the second period starts with a very reactive element sodium which is followed by the less reactive magnesium. Aluminium and silicon are less reactive than both sodium and magnesium. This is because the number of valence electrons increases, making it difficult to lose electrons. Moving further right in a period towards non-metals, the chemical reactivity gradually increases again.

Things to Know

Metallic character of a metal generally determines its level of reactivity.

Density

It is mass of a substance in a unit volume. Density of elements generally increases from top to bottom in a group but varies less significantly from left to right in a period. It is expressed in g/cm³.

Table (8.11) Densities of First Group Elements

First group	Li	Na	K	Rb	Cs
Elements density	0.53	0.97	0.89	1.63	1.879

Exercise

(1) Baruim (Ba) is present in 2nd group and 6th period. Answer the following questions about this element.
   (i) Is it a metal or a non-metal?
   (ii) Will it be electropositive or electronegative?
   (iii) What is the nature of its oxide?
   (iv) In which physical state you expect this element to exist?

(2) In which group and period you expect to find an element with the largest atomic radius?

(3) Can you predict the group number of the most electropositive and the most electronegative elements?

(4) Choose among the following the element having the lowest ionization energy and the element with highest electron affinity. Also assign its group number and period number Li, K, O, F, Cl.

(5) Which two elements of the periodic table react to give
   (i) A basic oxide and
   (ii) An acidic oxide?

Key Points

1. In the modern periodic table, elements have been arranged in the ascending order of their atomic numbers. There are eighteen groups and seven periods in this table.
2. In the modern periodic table, elements have been divided into s, p, d and f blocks.
3. Each period starts with an alkali metal and ends at a noble gas. A period also represents the completion of a shell.
4. The elements present in s and p blocks are called normal elements whereas those present in d and f blocks are called transition elements.
5. All elements in a group are expected to show similar chemical properties because of the same number of electrons present in their outermost shells.
6. Physical properties of the elements vary gradually as we move from left to right in a period and from top to bottom in a group.
7. Atomic sizes of elements decrease from left to right in a period and increase from top to bottom in a group.
8. Ionization energy values increase from left to right in a period and decrease from top to bottom in a group.
9. Electron affinity and electronegativity values increase from left to right in a period and decrease from top to bottom in a group.
10. Metallic character increases from top to bottom in a group and decreases from left to right in a period.

Exercise

1. Tick (✓) the correct answer.

(i) In which period and group will you place the element which is an important part of the solar cell?

(a) Third period and 14th group
(b) Second period and 14th group
(c) Third period and 15th group
(d) Third period and 16th group

(ii) Identify the electronic configuration of the outermost shell of a transition metal.

(a) ns^2np^4 (b) nd^xns^2
(c) ns^2np^6 (d) ns^2np^5

(iii) Which is the softest metal?

(a) Na (b) Ca (c) Al (d) Zn

(iv) A yellow solid element exists in allotropic forms which is also present in fossil fuels. Indicate the name.

(a) Carbon (b) Iodine
(c) Aluminium (d) Sulphur

(v) How many electrons can nitrogen accept in its outermost shell?

(a) 2 (b) 3 (c) 4 (d) 5

(vi) Which element is the most reactive element?

(a) Oxygen (b) Chlorine (c) Fluorine (d) Nitrogen

(vii) Which element has the highest melting point?

(a) Na (B) K (c) Rb (d) Cs

(viii) In what order does the metallic character change in the second group?

(a) Mg > Ca > Ba > Sr (b) Sr > Ba > Ca > Mg
(c) Mg > Sr> Ca > Ba (d) Ba > Sr > Ca > Mg

(ix) Which of the following best describe the correct order of oxygen, fluorine, and nitrogen's atomic radii?

(a) O < F < N (b) N < F < O (c) F < O < N (d) O < N < F

(x) The element having less value of ionization energy and less value of electron affinity is likely to belong to:

(a) Group 1 (b) Group 13 (c) Group 16 (d) Group 17

2. Questions for Short Answers

i. Why was atomic number chosen to arrange the elements in the periodic table?

ii. What is the significance of the word periodic?

iii. Why does the size of a period increase as we move down the periodic table?

iv. In a group, the elements have the same number of electrons in the outermost shell. Why is it so?

v. Do you expect calcium to be more reactive than sodium? Give the reason of your answer.

vi. Which element has the maximum atomic radius and which element has the minimum atomic radius in third period?

vii. Why are the most electronegative elements present in sixth and seventh groups?

viii. The first ionization energy value of magnesium is less than the second one. Give reason.

ix. Is it possible for two metals or two non-metals to form an ionic bond?

x. Which element has the least value of ionization energy and which element has the highest value of electronegativity?

3. Constructed Response Questions

i. Suppose a new element is discovered. Where would you like to accommodate this element in the periodic table?

ii. What is the first element of the periodic table? Will it lose an electron or gain it?

iii. Atomic radii of boron and aluminum are 88 pm and 125 pm respectively. Which element is expected to lose electron or electrons easily?

iv. How would you find the atomic radius of an atom?

v. Why is it not possible for oxygen atom to accept three electrons to form O^{3-} ion like nitrogen which can accept electrons to form N^{3-}?

4. Descriptive Questions

i. Which information is needed to locate the elements in the periodic table if you do not know its atomic number? Is atomic mass helpful for this purpose?

ii. How many blocks of elements are present in the periodic table? Are these blocks helpful in studying the properties of elements?

iii. Explain the variation in the following properties in the periods giving reasons.

(a) Atomic radius (b) Ionization energy

iv. Which physical properties of elements may lead us to know what type of bond it will form?

v. Write down the names of four non-metals which exist in solid state at normal temperature.

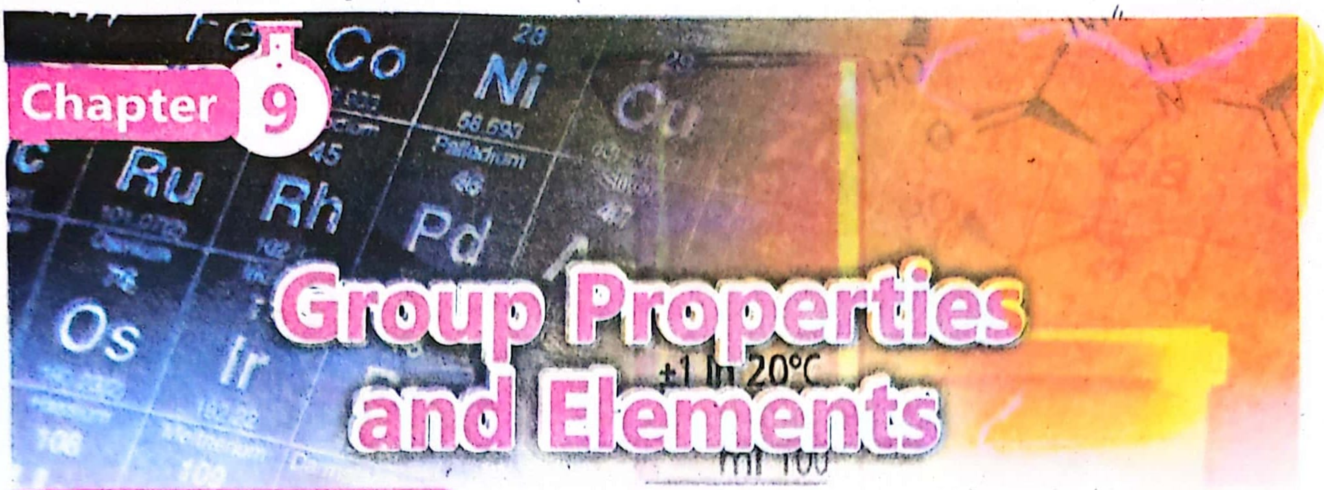
vi. Why do second and third periods have equal number of elements while all other periods contain different number of elements?

5. Investigative Questions

i. Arrangement of the elements in the form of a periodic table is a remarkable achievement of chemists. Comment on this statement citing the benefits of this table.

ii. Both lithium and beryllium show behaviour different from rest of the alkali and alkaline earth metals respectively. Can you think of the possible reasons for this difference?

iii. Modern periodic table is the amended form of the earlier table developed by Mandeleev. Elaborate how these two tables are different from each other.



Chapter 9 Group Properties and Elements

Student Learning Outcomes

After studying this chapter, students will be able to:

- Define group 1 Alkali metals as relatively soft metals with general trends down the group limited to decreasing melting point, increasing density and increasing reactivity.
- Predict properties of other elements in group 1, given information about the elements.
- Predict properties of elements in group 1 in order of reactivity given relevant information.
- Define group 17 halogens as diatomic non-metals with general trends limited to increasing density, and decreasing reactivity.
- Identify the appearance of halogens at rtp as fluorine as pale-yellow gas, chlorine as yellow -green gas, bromine as red -brown liquid, iodine as grey- black solid
- Explain the displacement reactions of halogens with other halide ions and also as reducing agents
- Predict the properties of elements in group 17, given information about the elements
- Analyse the relative thermal stabilities of the hydrogen halides and explain these in terms of bond strengths
- Transition elements
- Describe the transition elements as metals that: have high densities, high melting points, variable oxidation numbers, form coloured compounds and act as catalysts for industrial purposes. (Some examples include catalysts being used are the Haber process, catalytic converters, Contact process and manufacturing of margarine)
- Define the Group 18 noble gases as unreactive, monatomic gases
- Explain this in terms of electronic configuration properties of metals
- Compare the general physical properties of metals and non-metals (Specifically in terms of:

  a. thermal conductivity

  b. electrical conductivity

  c. malleability and ductility

  d. melting points and boiling points)

Elements present in a group of the periodic table show similar chemical properties owning to the presence of same number of electrons in their outermost shells. However, a small variation in the chemical properties of elements is expected because the atomic size increases down the group.

9.1 Properties of Group 1 Elements

All the elements present in group 1 have ns^1 configuration in their outermost shells. They are also called alkali metals. This single electron can be removed easily which makes these metals very reactive except the first element hydrogen which is a gas and a non-metal. When we move from top to bottom in this group, the atomic size increases. Owing to this, it becomes easier for elements to lose electron down the group which is reflected in the increased reactivities of the lower members of the group.

Lithium reacts with water steadily giving hydrogen and lithium hydroxide. Sodium reacts vigorously while potassium reacts violently with water giving their respective water soluble hydroxides.

$$2Li_{(s)} + 2H_2O_{(l)} \longrightarrow 2LiOH_{(aq)} + H_{2\,(g)}$$

$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2\,(g)}$$

$$2K_{(s)} + 2H_2O_{(l)} \longrightarrow 2KOH_{(aq)} + H_{2\,(g)}$$

Similarly, reaction of these metals with chlorine becomes more vigorous as we go down the group.

$$2Li_{(s)} + Cl_{2(g)} \xrightarrow{\text{Heat}} 2LiCl_{(s)}$$

$$2Na_{(s)} + Cl_{2(g)} \xrightarrow[\text{vigorous reaction}]{\text{Heat}} 2NaCl_{(s)}$$

$$2K_{(s)} + Cl_{2(g)} \xrightarrow[\text{violent reaction}]{\text{Heat}} 2KCl_{(s)}$$

Exercise

Keeping in view the trends of reactivity in first group elements how would they react with oxygen?

Interesting Information!

Li, Na and K are lighter than water but rubidium sinks in water. Cesium explodes on contact with water, possibly shattering the container.

Increase in the atomic size down the first group also weakens the interatomic attraction of the atomic metals. This fact makes them softer down the group and their melting points decrease.

As we go down the first group, both the size and volume of the atoms increase as the number of electrons and protons increases. But the increase in

mass of the elements is greater than the increase in volume; so the density which is defined as the mass per unit volume increases gradually down the group. Table (9.1)

Table (9.1) Physical Properties of First Group Metals

Metal	Li	Na	K	Rb	Cs
Melting point °C	180	98	64	39	28
Density g/cm^3	0.53	0.97	0.86	1.53	1.87

9.2 Properties of Group 17 Elements

All the elements in the group 17 have seven electrons in their outermost shells (n^2,s^5). They are electronegative non-metals because they have strong tendency to accept one electron to become an anion. They exist as diatomic molecules and behave as very reactive non-metals. Atomic radii, melting and boiling points of halogens increase when you go down the group. This is because the atoms get larger as they have more electrons. As a result the molecules which are formed from these atoms will attract each other with a stronger force. This is the reason why the melting and boiling points of bromine and iodine are higher than those of chlorine are fluorine.

Halogens react with alkali and alkaline earth metals to give salts. These elements are thus named as **Halogens** which means salt-forming elements. Unlike metals, the reactivity of halogens decreases from top to bottom in the group. This is due to the fact that atomic size increases down the group and tendency to accept electron from other atoms decreases making them less reactive.

Fluorine gas is pale yellow, chlorine gas yellowish green while bromine is fuming red-brown liquid. Iodine exists as shiny grey crystals which easily turn into dark purple vapours when they are warmed up.

Metal halides are formed when halogens react directly with alkali and alkaline earth metals. Metal halides behave usually as ionic compounds.

Oxidation is a process in which an electron is lost. The substance which loses an electron is called a reducing agent.

$$Na_{(s)} \longrightarrow Na^+ + e^-$$

On the contrary, reduction is a process in which an electron is gained and the substance which accepts the electron is called an oxidizing agent.

$$Cl_2 + 2e^- \longrightarrow 2Cl^-$$

Oxidation-reduction are simultaneous reactions. In other words an, electron is lost only when there is a substance to accept it.

Halogens are oxidising agents and their oxidising power decreases down the group.

$$Cl_2 \quad > \quad Br_2 \quad > \quad I_2$$

This fact gives a unique property to halogens when a halogen having more oxidising power displaces an ion of another halogen from its compound.

$$Cl_{2(aq)} + 2NaBr_{(aq)} \longrightarrow 2NaCl_{(aq)} + Br_{2(g)}$$

$$Cl_{2(aq)} + 2NaI_{(aq)} \longrightarrow 2NaCl_{(aq)} + I_{2(g)}$$

$$Br_{2(aq)} + 2NaI_{(aq)} \longrightarrow 2NaBr_{(aq)} + I_{2(g)}$$

Halogens react with hydrogen to give hydrogen halides. Hydrogen halides behave as strong acids in water.

$$H_2 + X_2 \longrightarrow 2HX$$

All halides exist in gaseous state at ordinary temperature except hydrogen fluoride which is a liquid. Bond length between hydrogen and halogen increases down the group because as the halogen atom gets bigger the bonding pairs of electron get further away from the halogen nucleus. The bond between hydrogen and halogen therefore gets weaker. The weaker the bond, the less heat energy it will need to break it. Hence the thermal stability of hydrogen halides decreases down the group.

$$HF \quad > \quad HCl \quad > \quad HBr \quad > \quad HI$$

Exercise

How do halogens react with water?

Interesting Information!

Water in swimming pool is sterilized with chlorine.

9.3 Group Properties of Transition elements

Elements present at the centre of the modern periodic table from group 3 to group 12 are called d block elements or transition elements. All transition elements are metals having the similar properties. Transition elements are often hard with higher densities. Their melting and boiling points are also high. These metals show variable oxidation states and the compounds they form are often coloured. They are malleable and ductile.

Transition metals and their compounds function as catalysts in many important chemical reactions. Metals often absorb other substances on their surface and activate them in this process. Iron, a transition metal, is used as a catalyst in one of the most important industrial reactions which gives ammonia. It is called Haber process. This ammonia is used to prepare urea fertilizer.

$$N_{2(g)} + 3H_{2(g)} \xrightarrow{Fe} 2NH_{3(g)}$$

Platinum was originally used as a catalyst in the contact process for the manufacture of sulphuric acid. This expensive catalyst is, however, rendered inactive due to the presence of arsenic as impurity in sulphur dioxide. Vanadium pentoxide (V_2O_5) is now preferred as a catalyst.

$$2SO_{2(g)} + O_{2(g)} \xrightarrow{V_2O_5} 2SO_{3(g)}$$

$$SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(aq)}$$

A catalytic converter is a device used in the exhaust of an automobile which converts more harmful gases produced in the engine to such gases which do not pollute the atmosphere. Platinum, palladium and rhodium are the catalysts used in catalyic converters.

A transition metal nickel is used as a catalyst for the hydrogenation of oils to give solid margarine. Margarine is less likely to spoil than butter.

Things to Know

Transition metals have high tensile strength. What does it mean?

9.4 Properties of Noble Gases

Elements present in group 18 of the modern periodic table are called noble elements. All noble elelments are monoatomic gases having very low boiling points, Helium (He), Neon (Ne) Argon (Ar), Krypton (Kr), Xenone (Xe) and Radon (Ra). All these gases have eight electrons (s^2p^6) in their outermost shells except He which has s^2 electronic configuration. Since their outer shells are complete, they show very little chemical reactivity.

9.5 Physical Properties of Metals and Non-metals

Metals and non-metals can be distinguished based on their physical and chemical properties.

Metals are defined as the elements which can generally form cations easily. They also tend to form metallic bond. Metals can be hammered into thin sheets. This property is called malleability. Metals can also be drawn into wires and this property is named as ductility. Metallic bonding in metals allows metals to be the best conductor of heat and electricity. Metals are lustrous which means that they have a shiny appearance. Due to high tensile strength metals can hold heavy weights. When metals are hit by an object, they make a ringing sound. Metals cannot be cut easily because they are hard substances.

Due to the presence of strong metallic bond metals generally have high melting and boiling points. Their densities are also very high. Alkali metals being soft metals are treated as exceptions. Example of metals include copper, silver, iron, aluminium, gold, platinum, zinc, etc.

Non-metals generally gain electrons easily. Non-metals show a greater variety of colours and physical states compared to metals. Non-metals cannot be beaten into thin sheets because being brittle they break into pieces when hammered. Sulphur and phosphorous exist in powdered form and cannot be made into sheets. Non-metals cannot be drawn into wires without breaking. Non-metals do not have free electrons due to which the bonds between their atoms are weak and they break down when stretched.

As there are no free electrons so non-metals cannot conduct heat and electricity. Graphite is the only exception. It conducts electricity because of its special crystalline arrangements. The electrons which are present between the layers of graphite crystal are loosely held and hence they can become mobile. The conduction of electricity of graphite is due to the mobility of these electrons.

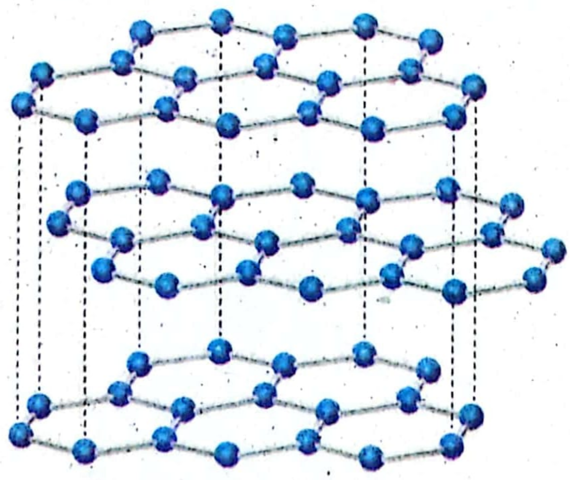


Fig (9.1): Graphite structure

Non-metals cannot be polished because they either exist in powder or gaseous form. Most of the powders are dull in texture. Due to non-ductile and non-malleable properties, non-metals are not strong at all. Their bonds being weak, break easily. All non-metals have low melting and boiling points. The melting point of sulphur is 115°C. Graphite and diamond have high melting points and these are exceptions. Non-metals have low densities as compared to metals. This means that in non-metals atoms are not strongly bound with each other. Examples of non-metals are oxygen, nitrogen, chlorine, sulphur, carbon and bromine, etc.

Interesting Information!

About 75% of all the elements in the periodic table are metals. There are total 20 non-metals which exist in solid or liquid or gaseous state at room temperature.

Comparison of the Physical Properties of Metals and Non-metals

Metals	Non-metals
1. Metals usually have high melting and boiling points.	1. Non-metals may be solids, liquids or gases at room temperature. They show wide range of melting and boiling points.
2. Metals are good conductors of heat and electricity.	2. Non-metals are poor conductors of heat and electricity (except graphite).
3. Metals can be made into different shapes by applying pressure. Metals can therefore be easily drawn into wires and sheets.	3. Non-metals are brittle.
4. Metals are usually lustrous solids (except mercury).	4. Non-metals are dull and cannot be polished (except iodine).
5. Metals are generally tough and strong.	5. Non-metals are neither tough nor strong.

Interesting Information!

According to one report nine elements are facing serious threat of extinction. Some of these elements are arsenic, gallium, gold, helium and zinc.

Key Points

1. Atomic size increases from top to bottom for elements of group-1. It becomes easier for these elements to lose their single electron.
2. Chemical reactivity of the elements in group-1 increases down the group.
3. Due to increase in their atomic sizes, the interatomic attraction decrease down the first group elements. This makes them softer and their melting points decrease down the group.
4. Due to increase in the size and volume of atoms, the densities of alkali metals increases down the group.
5. Group 17 elements or halogens exist as diatomic molecules. They are very reactive non-metals and react with alkali and alkaline earth metals.
6. Halogens are oxidising agents and their oxidising power decreases down the group.
7. In hydrogen halides the bond between hydrogen and halogen gets weaker as we go down the group. Hence, the thermal stability of hydrogen halides decreases down the group.
8. d – block elements or transition elements are all metals. They are often hard and have higher densities. These elements and their compounds are used as catalysts in important industrial reactions.
9. Noble gases are all unreactive gases because their outermost shells are complete.
10. Metals and non-metals have very different physical properties.

Exercise

1. Tick (✓) the correct answer.

(i) Which halogen will have the least reactivity with alkaline earth metals?

(a) Chlorine (b) Iodine
(c) Bromine (d) Fluorine

(ii) Which compound do you expect to be coloured?

(a) KCl (b) $BaCl_2$
(c) $AlCl_3$ (d) $NiCl_2$

(iii) In which element there exists the strongest forces of attraction between atoms?

(a) Mg (b) Ca
(c) Sr (d) Ba

(iv) Elements of which group are all coloured?

(a) Second group (b) Sixth group
(c) Fourth group (d) Fifth group

(v) Which halogen acid is unstable at room temperature?

(a) HBr (b) HI

(c) HCl (d) HF

(vi) Which oxide is the most basic oxide?

(a) Na_2O (b) Li_2O (c) MgO (d) CO

(vii) Which group elements are the most reactive elements?

(a) Transition metal group (b) First group

(c) Second group (d) Third group

(viii) The following solutions of a halogen and a sodium halide are mixed together. Which solution will turn dark because of a reaction?

(a) Br_2 and NaCl (b) Br_2 and NaF

(c) Cl_2 and NaF (d) Cl_2 and NaI

(ix) X is a monoatomic gas, which statement about this is correct?

(a) X burns in air (b) X is coloured

(c) X is unreactive (d) X will displace iodine from it

(x) Which property is correct for group 1 elements?

(a) Low catalytic activity (b) High density

(c) Low electrical conductivity (d) High melting point

2. Questions for Short Answers

i. Why does it become easier to cut an alkali metal when we move from top to bottom in a group I?

ii. Predict the reactivity of potassium towards halogens.

iii. In the following reaction, chlorine acts as an oxidising agent. Which is the reducing agent?

$$Cl_{2(aq)} + 2NaBr_{(aq)} \longrightarrow 2NaCl_{(aq)} + Br_{2(g)}$$

iv. Why does iodine exist in the solid state at room temperature?

v. How does Ni catalyse the reaction involving hydrogenation of oil?

3. Constructed Response Questions

i. Which noble gas should have the lowest boiling point and why?

ii. Compare the reactions of alkali metals with chlorine.

iii. Why are almost all the metals solids while non-metals generally exist as gases and solids?

iv. Name any three elements in the periodic table which exist as liquids.

v. Why are transition elements different from normal elements?

vi. Compare the reactivity of chlorine and bromine as an oxidising agent.

vii. Which element is the most reactive and which is the least reactive among halogens? Give two reasons to explain your answer.

4. Descriptive Questions

i. Explain the role of catalytic converter in an automobile.

ii. Why do the chemical reactivities of alkali metals increase down the group whereas they decrease down the group in case of halogens?

iii. Why are metals generally tough and strong whereas non-metals are neither tough nor strong?

iv. Both alkali metals and halogens are very reactive elements with roles opposite to each other. Explain.

v. Why hydrogen bromide is thermally unstable as compared to hydrogen chloride?

vi. Compare the properties of metals and non-metals.

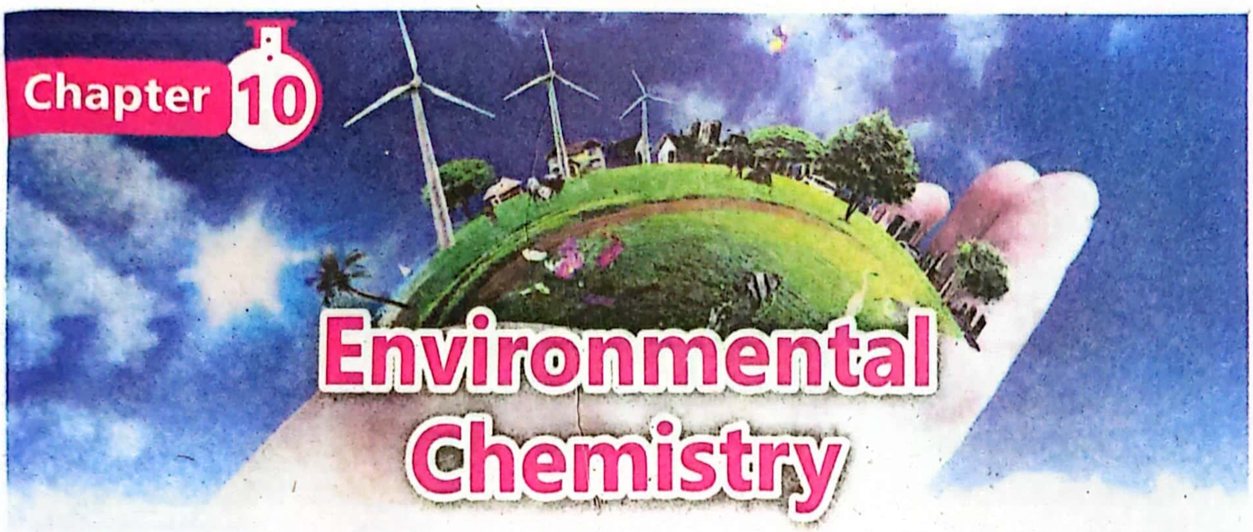
vii. V_2O_5 catalyst is preferred over platinum in the oxidation of sulphur dioxide. Give reasons.

5. Investigative Questions

i. Explain the role of sodium as heat transfer agent in the atomic nuclear power plant. Which property of sodium is utilized in this role.?

ii. Why and how does lithium behave differently from the rest of the alkali metals?

iii. Why aluminum metal is used in the manufacture of cooking utensils whereas magnesium is not considered useful for this purpose?



Chapter 10

Environmental Chemistry

Student Learning Outcomes

After studying this chapter, students will be able to:

- State that composition of clean, dry air is approximately 78% nitrogen, N_2, 21% oxygen, O_2, and the remainder as a mixture of noble gases and carbon dioxide, CO_2
- State the major sources of air pollutants (Some examples include:
  a. carbon dioxide from the complete combustion of carbon-containing fuels
  b. carbon monoxide and particulates from the incomplete combustion of carbon-containing fuels
  c. methane from the decomposition of vegetation and waste gases from digestion in animals
  d. oxides of nitrogen from car engines
  e. sulfur dioxide from the combustion of fossil fuels which contain sulfur compounds
  f. ground level ozone from reactions of oxides of nitrogen, from car engines, and volatile organic compounds, in presence of light)
- State the adverse effects of air pollutants (Some examples include:
  a. carbon dioxide: higher levels of carbon dioxide leading to increased global warming, which leads to climate change
  b. carbon monoxide: toxic gas
  c. particulates: increased risk of respiratory problems and cancer
  d. methane: higher levels of methane leading to increased global warming, which leads to climate change
  e. oxides of nitrogen: acid rain, photochemical smog and respiratory problems
  f. sulfur dioxide: acid rain and haze)
- Explain how the greenhouse gases carbon dioxide and methane cause global warming, (Some examples include:
  a. the absorption, reflection and emission of thermal energy
  b. reducing thermal energy loss to space

- Describe the role of sulfur in the formation of acid rain and impact on the environment. Describe the strategies to reduce the effects of major environmental issues (some examples include:

  a. climate change: planting trees, reduction in livestock farming, decreasing use of fossil fuels, increasing use of hydrogen and renewable energy, e.g. wind, solar.

  b. Acid rain: Use of catalytic converters in vehicles, reducing emissions of sulfur dioxide by using low sulfur fuels and flue gas desulfurization with calcium oxide).
- Describe the role of NO and NO_2(subscript) in the formation of acid rain, both directly and through their catalytic role in the oxidation of atmospheric sulfur dioxide.
- Explain how oxide of nitrogen form in car engines and describe their removal by catalytic converters, e.g. $CO + 2NO \rightarrow 2CO_2 + N_2$
- Define photosynthesis as the reaction between carbon dioxide and water to produce glucose and oxygen in the presence of chlorophyll and using energy from light.
- Analyze how to use tools to reduce personal exposure to harmful pollutants (some examples include the usage of masks, air quality indices and CO detectors).
- Identify high risk situation in life including those where long-term exposure to these pollutants can lead to respiratory issues and reduction in quality and longevity of life.

Introduction

The branch of Chemistry which deals with the study of chemicals and other pollutants in the environment is called **environmental chemistry**. It also covers the adverse effects of these chemicals on living and non-living things.

Environmental chemistry is a part of environmental education. The objective of which is to enlighten the people particularly the students, about the importance of protection and conservation of our environment. The need for this environmental education, both formal and non-formal, is keenly felt at the national level.

Since the start of industrial revolution, human activities have played havoc with the atmosphere of the Earth. The gases which are released due to the increasing use of fossil fuels (natural gas, coal and petroleum) have polluted the atmosphere upto such an extent that it is difficult to breathe air in some areas of metropolitan cities.

The excessive use of fertilizers, insecticides and pesticides, etc. for agriculture purposes has proved to be harmful for animals, birds and human beings. The situation is turning serious for every passing day and there is an urgent need to control the emission of pollutants to the atmosphere.

10.1 Composition Of Atmosphere

Earth is covered with a blanket of air called the "atmosphere" which is made up of several layers of gases. Air is essential for life on Earth, for animals to breathe and for plants to make their food. It contains more nitrogen than any other gas.

The components of the atmosphere may be divided into major and minor constituents. The amount of these different gases in the air varies slightly from place to place, season to season and day to night. The percentages of these constituents by volume are given in Table (10.1).

Table (10.1): Major and Minor Constituents of the Atmosphere

Major Constituent	Percentage	Minor Constituent	Percentage
Nitrogen	78.0	Carbon dioxide	0.04
Oxygen	21.0	Noble gases	About 1.0
Argon	0.934	Water Vapours	Variable Depending upon the humidity

Exercise

1. At which time of the day and night you expect humidity to be maximum?
2. Which gas is released when carbonated drinks open?

Interesting Information!

Environmental science helps us to understand the complex interactions that occur in our ecosystems and the impacts on human life.

10.2 Air Pollutants

Any substance (solid, liquid or gas) in the air which has adverse effect on human health and quality of life is called an air pollutant.

The concentration of a pollutant is expressed in parts per million (ppm). A concentration of one ppm means one part of pollutant per million part of solid, liquid or gas mixture in which the pollutant is formed.

Every Individual should try to.....

Pour liquid waste into sewers not in open drains, river and sea. Stop using environmentally hazardous substances.

Major Air Pollutants

Air is not always clean as it should be. There are seven types of harmful substances which account for more than 90% of air pollution. These are fast growing sources of air pollution created by our day-to-day activities. The detail of these substances is mentioned below.

(i)	Carbon dioxide	(CO_2)
(ii)	Carbon monoxide	(CO)
(iii)	Oxides of nitrogen	(NO, NO_2) collectively referred to as NO_x
(iv)	Oxides of sulphur	(SO_2, SO_3) collectively referred to as SO_x
(v)	Hydrocarbons	(Methane, ethane, etc.)
(vi)	Particulates	(Dust, pollens, metallic compounds)
(vii)	Ozone	(O_3)

Sources of Air Pollutants

Millions of tonnes of pollutants are emitted into the atmosphere each year as a result of human activities. The major activity among them is the complete and incomplete combustion of fossil fuels which alone is responsible for most of our pollution problem.

Burning of fossil fuels (oil natural gas, coal) produce carbondioxide, carbon monoxide, NO_x, SO_x, CH_4, ash, smoke and suspended particles.

Many of these pollutants are also released into the air by natural processes e.g., volcanic eruption releases large quantities of CO_2, SO_2 and particulates. Methane is released in the air by the decomposition of vegetation. It is also present in waste gases produced during digestion in animals.

Rapid growth of population, urbanization, industrialization and transportation are the main factors which are responsible for environmental pollution. All these factors are increasing in every city of the world especially in the last half century. These pollutants are affecting the environment very badly.

Another pollutant ozone (O_3) is formed when heat and sunlight cause chemical reaction between oxides of nitrogen (NO_x) and volatile organic compounds (hydrocarbons).

In winter the smoke present in the atmosphere is mixed with fog to form what is called **smog**. Many cities of Pakistan are completely covered with the blanket of smog. Its suspension is caused by a combination of factors which

include industrial pollution, vehicle emission and crop burning. These factors are responsible for the accumulation of nitrogen oxides, sulphur dioxide, particulate matter and volatile organic compounds in air.

The following Table (10.2) shows the major air pollutants and their harmful effects on human beings and on the environment.

Table (10.2) Pollutants and Their Harmful Effects

No.	Pollutants	Harmful Effects
1	Carbon dioxide (CO_2)	Higher levels of carbon dioxide lead to increased global warming which can cause ice caps to melt and oceans to warm, causing sea levels to rise. Extreme weather changes such as heat waves, heavy rains and wild fires also occur.
2	Carbon monoxide (CO)	It is extremely poisonous gas that can cause suffocation and death. Carbon monoxide is very toxic gas that stops the red blood cells in animal's blood from carrying oxygen that body needs.
3	Oxides of nitrogen (NO_x) NO, NO_2	NO_x can damage lungs; irritate the eyes and damage vegetation. It can also cause acid rain which affects buildings and statues made of limestone.
4	Oxides of sulphur (SO_x) SO_2, SO_3	SO_x irritates the eyes and causes breathing difficulties and acid rain.
5	Hydrocarbons	They can cause pneumonia, coughing, many other breathing and lung diseases. They also cause global warming.
6	Particulate matter	Irritates the eyes and can also cause severe breathing problems for people with asthma. It also makes clothes dirty. Also, visibility is reduced because it produces haze in the air.

7	Ozone	Breathing ozone can cause a variety of health problems including chest pain, coughing, throat irritation and congestion.

Apart from the effects of these pollutants, the smog present in the atmosphere can lead to health complications like allergies, asthma and lung infections. It also inhibits the growth of plants by reducing the amount of carbon dioxide absorbed during photosynthesis.

Amazing Facts about the Environment

1. 78% of marine mammals are at risk of choking with plastic.
2. Humans can use only 1% of all available water.
3. The world has over 3.04 trillion trees. 27000 out of them are cut down daily to make toilet paper.

Exercise

How does air pollution affect plants?

10.3 Acid Rain

When rain water has pH between 4.2 and 4.4, it is known as acid rain. In 1852, Robert Angus Smith was the first to show the relationship between acid rain and atmospheric pollution in Manchester, England. He is sometimes referred to as the "Father of Acid Rain". Burning of fossil fuels releases harmful gases into the atmosphere. These gases (SO_2, SO_3) are produced due to the presence of sulphur in the fossil fuels. SO_2 is converted to SO_3 in the presence of oxides of nitrogen of the atmosphere. Oxides of nitrogen are produced mostly by the direct combination of atmospheric oxygen and nitrogen in the industrial and domestic combustion processes. They are also produced by the combination of atmospheric nitrogen and oxygen in the presence of lightning. Significant amount of nitrogen oxides is produced by the reactions taking place in automobile engines Fig.(10.1).

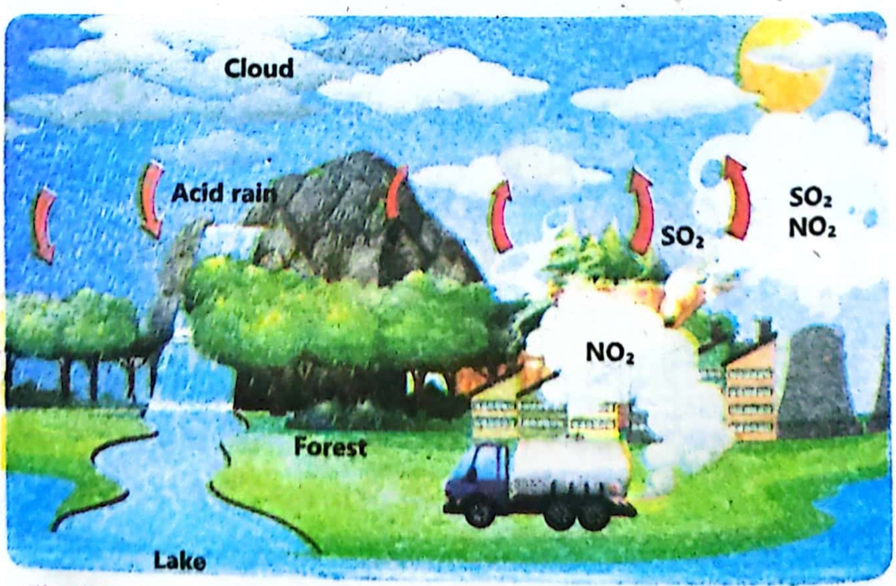


Fig (10.1): Acid rain usally falls far from the site where the acidic oxides are generated.

$$S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)}$$

$$2SO_{2(g)} + O_{2(g)} \xrightarrow{\text{Oxides of nitrogen}} 2SO_{3(g)}$$

$$SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(aq)}$$

$$N_{2(g)} + O_{2(g)} \longrightarrow 2NO_{(g)}$$

$$2NO_{(g)} + O_{2(g)} \longrightarrow 2NO_2$$

$$2NO_{2(g)} + H_2O_{(l)} \longrightarrow HNO_{3(g)} + HNO_{2(g)}$$

These gases mix with the moisture that is always present in the air to form acid droplets. Wind can carry these acidic droplets to huge distance.

Finally, these droplets return to the ground as acid rain, acid hail, snow and even fog. Acid rain looks, feels and tastes like clean rain. Its corrosive nature causes widespread damage to the environment.

Interesting Information!

There is a giant floating patch of garbage spread over in the Pacific Ocean. It contains about 100 million tons of garbage.

Effects of Acid Rain

Acid rain causes a number of adverse effects. It tends to increase acidity of the soil, threatens humans and aquatic animals, destroys forests, and reduces agricultural productivity. Besides, it can corrode buildings, monuments, statues, bridges and railings. Most important adverse effects of acid rain are as follows:

(i) Soil

Acid rain makes soil more acidic. It dissolves and washes away nutrients present in the soil which are needed by plants. It can also dissolve toxic substances such as aluminium and mercury which are naturally present in the soil.

(ii) Plants

Acid rain can damage vegetation and plants. Many plants cannot live or grow in acidic soil. Tree roots hold the soil together on hills and mountain areas. If the trees are destroyed then the soil is washed away and new plants cannot grow there.

(iii) Aquatic Life

Acid rain falls into drains, streams, lakes, marshes, rivers and damages the aquatic life. Acid rain can make water too acidic for animals to live in. Due to this, many lakes and rivers no longer have fish.

(iv) Human Health

The acidification of surface water does not affect life directly. However, toxic substances leached from the soil can pollute land water supplies and damage human health.

(v) Agriculture

Crops are less affected by the acid rain than forests. Farmers can prevent acid rain damage by monitoring the conditions of the soil and when necessary adding crushed lime (CaO) to neutralize the acid.

(vi) Human-made Structure

Acid rain and the dry deposition of acidic particles damage buildings, statues, automobiles, and other structures made of stone, metal etc. Historical buildings like Parthenon in Athens (Greece) and the Taj Mahal in Agra (India) are deteriorating due to acid rain.

Exercise

1. Which acids are made when SO_2 and NO_2 dissolve in rain?
2. What happens to the soil if trees are destroyed by acid rain?

10.4 Global Warming (Greenhouse Effect)

The progressive warming up of the Earth's surface due to blanketing effect of man-made carbon dioxide, methane, water vapours and other gases in the atmosphere is called Greenhouse Effect.

The sun emits short wave radiation that passes through greenhouse gases to heat the surface of the Earth. At night, the hot Earth's surface emits longwave radiation that is mostly absorbed by green house gases. This process of absorption prevents the radiation to reach space, reducing the speed at which the Earth can cool off (Fig 10.2). This increases the temperature of the Earth and causes global warming.

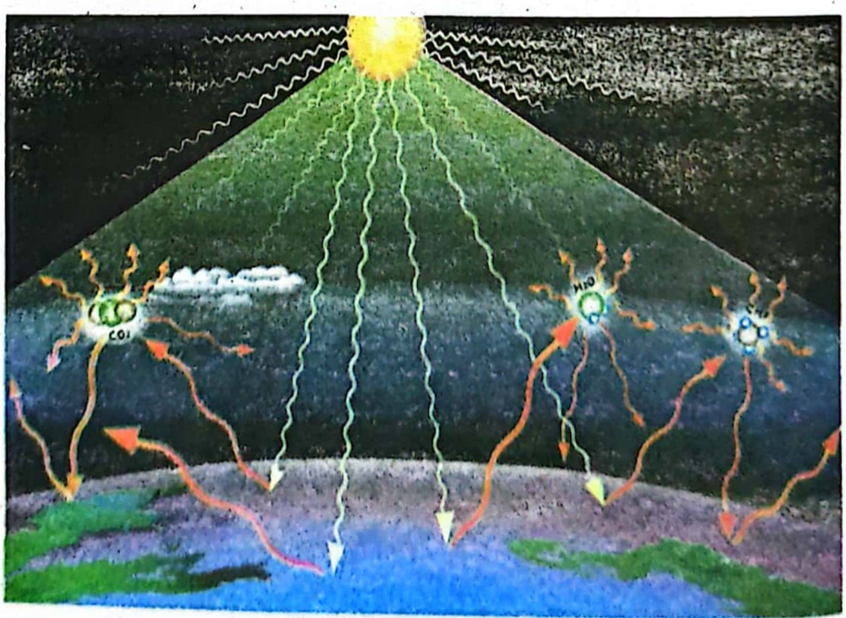


Fig 10.3: Global warming/greenhouse effect

Higher the concentration of carbon dioxide and other greenhouse gases, greater will be the absorption of thermal radiation and greater will be the increase in global warming.

Do you Know?

Greenhouse effect, depletion of ozone and acid rain are the global effects of pollution.

Sources of Greenhouse Gases

Due to the burning of the large amount of coal, oil and natural gas, the amount of green house gases, carbon dioxide together with other gases in the atmosphere has increased for the last 300 years. About half of this carbon dioxide is utilized by plant life during photosynthesis. As human beings cut down forests, the capacity of the trees to remove CO_2 from the air is diminished.

Methane is another green house gas which causes adverse effects. The increase in its concentration in air is due to the increased decay of vegetation matter, digestion in animals and increased farming of the rice fields.

Sunlight passes through the greenhouse gases and warms everything on the Earth

The Earth warms up and gives out heat. Some heat passes through the greenhouse gases but some is trapped inside, warming up the Earth.

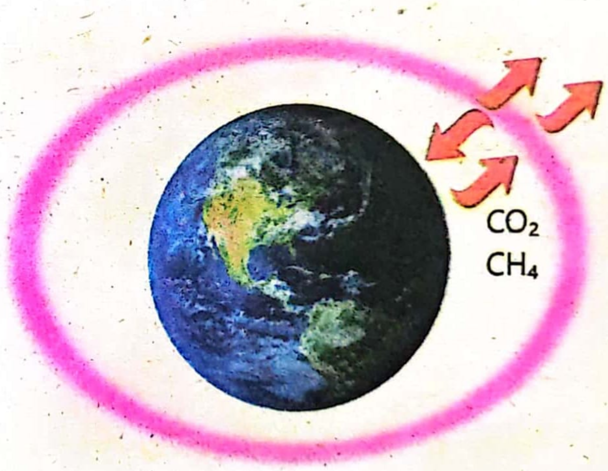


Fig (10.3): Greenhouse gases

Interesting Information!

Water vapours are Earth's most abundant greenhouse gas. It is responsible for about half of Earth's greenhouse effect.

The remaining carbon dioxide which is not utilized in photosynthesis goes on accumulating in the lower areas of the atmosphere and forms a thick dense layer. This layer behaves like glass sheet of greenhouse that allows the incoming solar radiation but does not allow it to escape outside; as a result of this the average temperature of the Earth rises.

A rise of a few degrees in temperature may seem small, but it can be enough to cause significant changes in the climate. At the moment, it is difficult for scientists to say how big the changes will be and where the worse effects will occur. This can damage agriculture and food production as well.

Exercise

1. How do living things add and plants remove carbon dioxide from the air?
2. Which gas do rice plants produce?
3. Which gas is given out by rotting garbage?

The effects of climate change may be physical, ecological, social or economical. Following are four adverse effects of the global warming.

Rise in Sea level

Higher temperature will make the water of the seas and oceans expand. Ice melting in the Antarctic and Greenland will flow into the sea and it results in higher sea levels. This phenomenon will threaten the low-lying coastal areas of the world such as the Netherlands and Bangladesh Fig (10.4).

Fig (10.4): Global warming at sea level

Increase in Rainfall

There may be enormous increase in rainfall in a few regions which may increase the sea level. This ultimately will cause worldwide floods endangering survival of living species. Fig (10.5)

Fig (10.5): Effects of Climate Change

Effects on Agriculture

The changes in the weather will affect the types of crops grown in different parts of the world. Some crops such as wheat and rice grow better in higher temperature but other plants such as maize and sugar cane do not Fig (10.6).

Fig (10.6): Global warming risking the agriculture drought.

Hot Summer and Winter

In moderate region, the winter will be shorter and warmer and the summer will be longer and hotter Fig (10.7).

Fig (10.7): Summer and winter

Interesting Information!

Since 1990, we have lost around 28 trillion tonnes of ice. At presents, its melting rate is 1.2 trillion tonnes per year.

10.5 Strategies to Reduce Environmental Issues

Huge amount of pollutant gases are thrown out in the atmosphere by burning fossil fuels. Automobiles, aeroplanes, industrial machines and coal-fired electricity generating plants, etc. are mainly responsible for the extremely polluted air especially in big cities. Scientists have developed a number of different ways to control this menace of pollution.

Planting trees is thought to be very helpful in removing the air pollution. A well-known process carried out by plants is photosynthesis in which plants clean the air through absorption of carbon dioxide and releasing oxygen. This famous reaction takes place in the presence of sunlight and it is catalysed by chlorophyll, the green pigments present in leaves.

$$6\,CO_{2\,(g)} + 6H_2O_{(l)} \xrightarrow[\text{Sunlight}]{\text{Chlorophyll}} C_6H_{12}O_{6(s)} + 6O_{2\,(g)}$$

The particulate matter present in the atmosphere is also removed by plants when it deposits on leaves, branches and trunk surfaces.

Catalytic converters are used in the exhaust system of modern-day automobiles to reduce the emissions from an internal combustion engine. Due to non-availability of enough oxygen the carbon fuel in engine does not burn completely into carbon dioxide and water. Thus, toxic by-products like CO and hydrocarbons are produced.

A three-way catalytic converter performed the following three functions simultaneously.

1. It reduces nitrogen oxides into elemental nitrogen and oxygen.

$$2NO_{x\,(g)} \longrightarrow N_{2(g)} + xO_{2\,(g)}$$

2. It oxidizes CO to CO_2

$$2CO_{(g)} + O_{2\,(g)} \longrightarrow 2CO_{2\,(g)}$$

3. It oxidizes hydrocarbons into CO_2 and H_2O

$$C_xH_{4x\,(g)} + 2xO_{2(g)} \longrightarrow x\,CO_{2(g)} + 2x\,H_2O_{(l)}$$

Similarly, the emission of sulphur dioxide can be decreased either by using fuels which have significantly less sulphur contents or using flue gas desulphurization process. This process can remove sulphur dioxide gas from the exhaust gases of fossil fuel. Flue gas is the mixture of waste gases produced when fossil fuels are burnt in power plants. The desulphurization process involves the addition of adsorbents like calcium oxide which can remove upto 95% of the sulphur dioxide from the flue gas.

To discourage the excessive use of fossil fuels in our daily lives it is urgently required to use the renewable resources to meet our energy needs. Renewable resources are those resources that can continue to exist despite being consumed over a period of time even as they are used.

These resources include Sun, wind, water, geothermal and biomass. Solar energy and wind energy have been proved to be very effective ways of generating electricity without damaging the environment.

Interesting Information!

Radon is a natural radioactive gas. When it builds up indoors to high level, it increases the rise of lung cancer.

How to avoid harmful effects of air pollution?

Air quality index (AQI) is a rating system that shows how bad is the atmosphere around you. An (AQI) value under 50 is considered good in quality. This means that it is safe for you to spend time outdoors without posing a risk to your health. As the (AQI) number increases, so does the risk to health. An (AQI) over 300 is considered hazardous. Children under 18, adults over 65, people with chronic heart or lung disease and diabetic people are high risk groups. Outdoor workers can also be considered at higher risk because of prolonged exposure.

The following steps should be taken when air quality is bad.

1. Reduce the time you spend outdoors. Also reduce the intensity of outdoor activity. According to experts the chances of being affected by unhealthy atmosphere increase if a person stays outdoor for longer periods or involve in more laborious activity outdoor.
2. If you are forced to go out, then consider wearing a mask. Unfortunately, not all the masks provide adequate safety against particulate matter. Cloth or dust mask may not effectively filter out the finer particles. However, well fitted N95 masks have better filtration capabilities and may be safer to use.
3. Keep your indoors healthy by keeping the windows and doors closed. If it is difficult to maintain clean air in the entire room then create a clean room by switching on air conditioner or air cleaner.
4. If you experience such symptoms that worry you, talk to your doctor.
5. Install carbon monoxide detector to detect the increased level of carbon monoxide. These higher levels of CO may occur due to faulty fuel burning appliances.

Breathing in polluted air by these high risk groups may affect their lungs, heart and brain. Air pollutants can enter their blood stream and can cause coughing or itching of eyes which may lead to poor quality of life, hospitalization, cancer or even premature death.

Key Points

1. Earth is covered with a blanket of air called atmosphere which is made up of several layers of gases.
2. Any substance in the air which has an adverse effect on human health, quality of life and natural functioning of the ecosystem is called a pollutant.
3. Oxides of carbon, nitrogen and sulphur alongwith methane and particulate matter are the main pollutants in the air.
4. Major sources of pollutants are due to human activities especially the burning of fossil fuels.
5. Pollutants have extremely adverse and harmful effects not only on the human beings but on the whole ecosystem.
6. Acid rain is formed due to the presence of oxides of nitrogen and sulphur when they are mixed with the moisture present in the atmosphere.
7. The progressive warming of Earth's surface due to blanketing effect of CO_2, CH_4 and other gases present in the atmosphere is called greenhouse effect. This effect has increased the temperature of the Earth.
8. Every effort should be made to reduce the harmful effects of the pollutants. These efforts include discouraging the use of fossil fuels, planting trees and using the renewable sources of energy.
9. Steps should be taken to avoid the harmful effects of pollution especially on the people who are at higher risk.

Exercise

1. Tick (✓) the correct answer.

(i) Which gases are responsible for greenhouse effect?

(a) SO_2, NO_2 (b) NO_2, CO
(c) CO_2, CH_4 (d) O_2, N_2

(ii) Indicate the source of sulphur which is responsible for the presence of oxides of sulphur in the atmosphere.

(a) Decomposition of vegetation
(b) Waste gases from digestion of animals
(c) Photochemical smog
(d) Combustion of fossil fuels

(iii) Concentration of which gas in the atmosphere is decreased by photosynthesis in plants?

(a) Oxygen (b) Nitrogen
(c) Carbon dioxide (d) Water vapours

(iv) Which substance remains unaffected in the catalytic converter fixed in the exhaust of the automobiles?

(a) CO_2 (b) CO

(c) NO (d) No_2

(v) People of which age groups are most affected by the air pollution?

(a) Young adults

(b) Middle age people

(c) Children

(d) Both children and aged people

(vi) In which area there is a greater possibility of acid rain?

(a) Around village (b) Around big cities

(c) Around industrial area (d) Around water bodies

(vii) Why is smog not felt in summer?

(a) Because fog is not present in summer

(b) Because due to heat of the Earth the smoke rises up

(c) Because in summer smoke and fog cannot mix with each other

(d) Because less fossil fuels are burnt in summer

(viii) Which catalyst is used in the catalytic converter fixed in the exhaust systems of automobiles?

(a) Ni (b) Cu (c) Pt, Pd and Rh (d) CaO

(ix) Which components are essential for the formation of photochemical smog?

(a) CO, NO_2, CO_2

(b) NO_2, volatile organic compounds, sunlight

(c) CO_2, NO_2, sunlight

(d) Volatile organic compounds, NO_2, CO

(x) Which gases contribute towards the formation of acid rain?

(a) Oxides of carbon (b) Oxides of sulphur

(c) Oxides of nitrogen (d) Both the oxides of nitrogen and sulfur

2. Questions for Short Answers

i. What is the main objective of environmental education?

ii. How is particulate matter released in the atmosphere?

iii. Which gas is more poisonous, CO_2 or CO?

iv. How does acid rain affect forests?

v. In what way sulphur present in fossil fuels becomes dangerous?

vi. Name any three major sources responsible for the greenhouse effect.

vii. How is wind energy useful for us?

3. Constructed Response Questions

i. How is the excessive use of insecticides and pesticides harmful for birds?

ii. Percentage of CO_2 in air is only 0.04%. Then how does it become harmful for the ecosystem?

iii. Why only some pollutant gases present in the atmosphere cause green house effect while others do not?

iv. How can you reduce the emission of CO present in the gases emitted by the burning of fuel in the automobile engines?

v. Mention three different ways in which solar energy can be usedful for us.

4. Descriptive Questions

i. Describe the harmful effects of the major pollutants present in the air.

ii. Explain greenhouse effect. How is global warming dangerous for us?

iii. What is air quality index? What information does it convey?

iv. Who are high risk groups and why is pollution more dangerous for them?

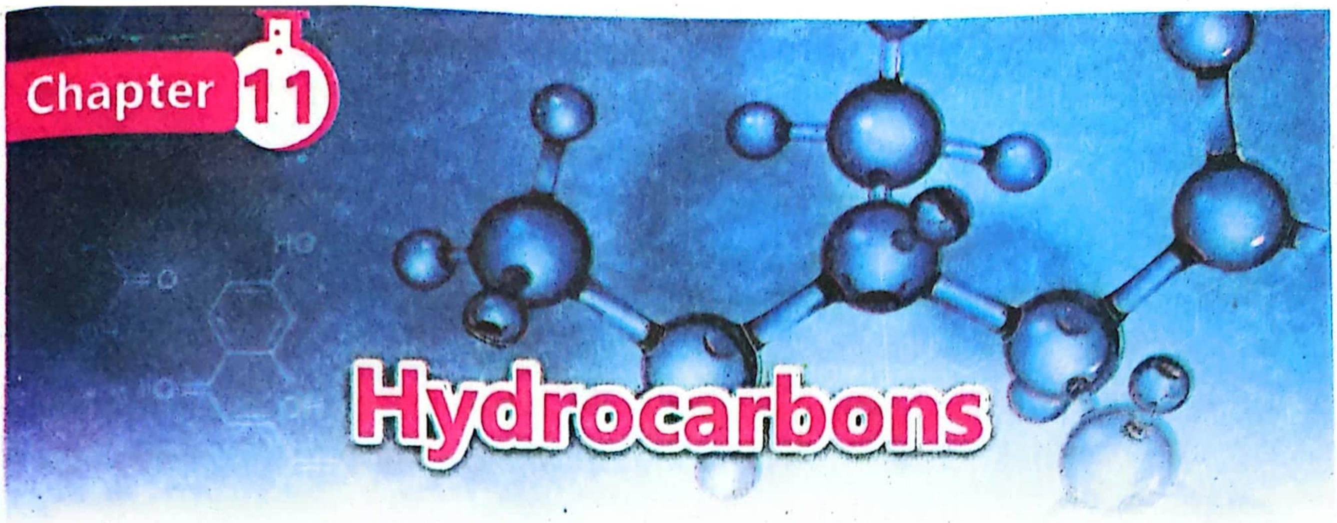
v. Describe three strategies to address environmental issues.

5. Investigative Questions

i. Major Pakistani cities experience a very high AQI in winter which renders them the most polluted cities in the world. Point out some of the major causes of high AQI in these cities.

ii. Why does AQI not rise in Pakistan in hot days of summer?

iii. How has climate change affected Pakistan during the last five years?



Chapter 11

Hydrocarbons

Student Learning Outcomes

After studying this chapter, students will be able to:

- Describe the properties of alkanes as being generally unreactive, except in terms of combustion and substitution by chlorine
- State that in a substitution reaction one atom or group of atoms is replaced by another atom or group of atoms
- Describe the substitution reaction of alkanes with chlorine as a photochemical reaction, and draw the structural or displayed formulae of the products.
- Describe, using symbol equations, preparation of alkanes from cracking of larger hydrocarbons, hydrogenation of alkenes and alkynes, and reduction of alkyl halides

Introduction

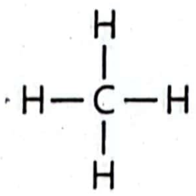
All the organic compounds are known to contain carbon as an essential element. This fact has led us to define organic chemistry as the chemistry of carbon compounds. The ionic compounds like carbonates, cyanides, carbides and cyanates, etc. and the oxides of carbon are, however, classified as inorganic compounds. Apart from carbon, most of the organic compounds contain hydrogen and oxygen as well.

Organic compounds are famous for their large number and diverse behaviour. Several million organic compounds are known to exist naturally or have been synthesized in the laboratory. Organic molecules are usually large and more complex in nature. They include biomolecules like proteins, enzymes, carbohydrates, lipids, vitamins and nucleic acids, pharmaceuticals and synthetic fibres, etc.

The number of compounds formed by the element carbon is far more than the total number of compounds formed by all the rest of elements put together. This is due to some unique properties of carbon.

The element carbon is present at the center of the periodic table and it is generally not possible for it to gain or lose electrons to form ionic bonds. Therefore, it forms four covalent bonds. Because of it small size, these covalent bonds are short and strong enabling carbon to give strong and stable bond with itself and with hydrogen, oxygen and nitrogen. The self-linking property of carbon is called **catenation** and due to this it forms long, straight and branched chains and rings.

Carbon atom mostly forms covalent bonds with other carbon, hydrogen, oxygen, nitrogen and halogens. Compounds in which carbon forms single bond with other atoms are called saturated compounds. These compounds are generally represented by their structural formulas. Methane a saturated compound, is represented by the following structural formula.

 **or its condensed form is CH_4**

A structural formula thus shows the symbols for the atoms present in the compounds connected by short lines which represent the bonds present in them. Other examples of saturated compounds are C_2H_6, CH_3Cl, CH_3OH, CH_3NH_2, etc.

Exercise

What do you understand by the term structural formula of an organic compound?

11.1 Hyrocarbons

The family of hydrocarbons constitutes a very large number of simple organic compounds containing carbon and hydrogen only. Most of the fuels which we use every day, for example, natural gas, LPG (Liquefied Petroleum Gas), CNG (Compressed Natural Gas), petrol, diesel and kerosene oil, are all simple hydrocarbons. These hydrocarbons also serve as a feedstock to prepare more complex and useful compounds like plastics, medicines, synthetic fibres, paints and varnishes.

Hydrocarbons are classified into several structural types called, alkanes, alkenes, alkynes and aromatic hydrocarbons. Only alkanes are discussed in this chapter.

Methane is the simplest alkane and it is mainly used as a fuel but it is also used to make hydrogen gas, carbon black, carbon disulphide, chloroform and hydrocyanic acid, etc.

11.2 Alkanes

Alkanes are the simplest family of hydrocarbons with only carbon – carbon and carbon – hydrogen single bond. Alkanes are also called saturated hydrocarbons because all the four valencies of carbon in them are fully utilized either by hydrogen atoms or by other carbon atoms. For example, in ethane (CH_3-CH_3) both the carbon atoms are fully saturated with single bonds.

Alkanes are represented by a general formula C_nH_{2n+2} (where n is an integer). Methane being the simplest hydrocarbon, is also called parent hydrocarbon.

As a result of the great complexity and large number of organic compounds, it is not possible to name each and every compound individually. The International Union of Pure & Applied Chemistry has devised a systematic way of naming organic compounds called IUPAC nomenclature.

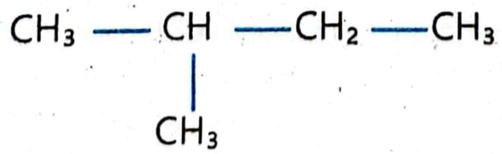
According to IUPAC system of nomenclature, the entire name of an organic compound has three parts:

1. Root: It tells us the number of carbon atoms in the longest continuous chain present in the molecule. The roots upto ten carbon atoms are shown in Table (11.1).
2. Suffix: It is added after the root and tells us about the class of organic compounds.
3. Prefix: It is indicated before the root and tells us about the group or groups attached to the longest chain.

Table (11.1)

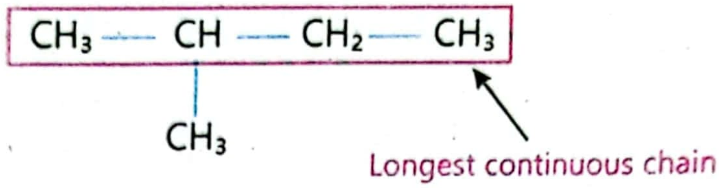
Root	No. of Carbon atoms
Meth-	1
Eth-	2
Prop-	3
But-	4
Pent-	5
Hex-	6
Hept-	7
Oct-	8
Non-	9
Dec-	10

To explain the above system, let us name the following compound.

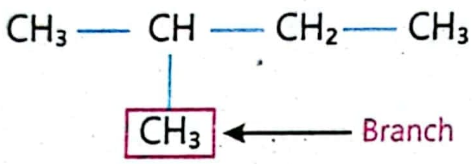


(a) Identify the longest continuous chain present in the compound.
(b) Identify the class of organic compounds.
(c) Identify the substituent or substituents if present.

This organic compound contains four carbon atoms in the longest continuous chain and it belongs to the family of organic compounds called Alkanes. The root is therefore But- and the suffix –ane added to this. The organic compound will thus be given the name Butane.

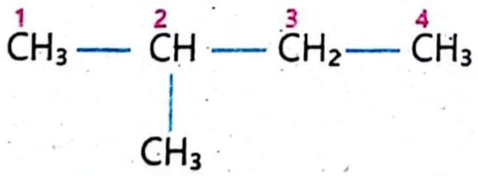


The name of the only branch methyl- will be added to this name as prefix. So the name will become:



Methylbutane

To specify where the branch occurs, the longest continuous chain is numbered starting from the end closest to the branch. This number is then attached to this prefix. The name of the above compound will then be:


2-Methylbutane or iso-Butane

If butane has no branch it is called normal or n-butane.

$$CH_3 - CH_2 - CH_2 - CH_3$$

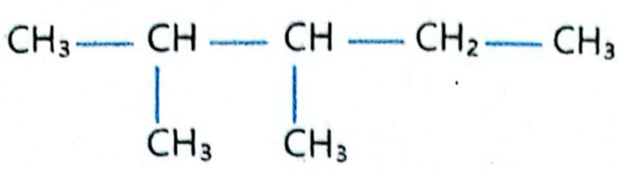
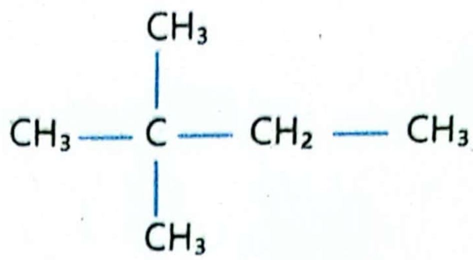
n-Butane

Interesting Information!

The distinguishing feature of alkanes making them distinct from other compounds is their lack of reactivity towards usual chemical reagents.

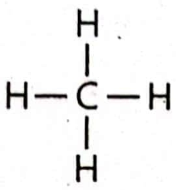
Exercise

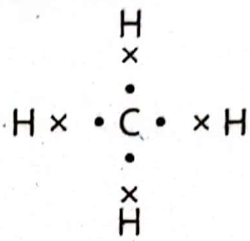
Name the following compounds according to IUPAC system of nomenclature.

(i)  (ii) 

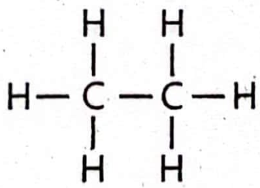
Electron cross and dot structures of Alkanes

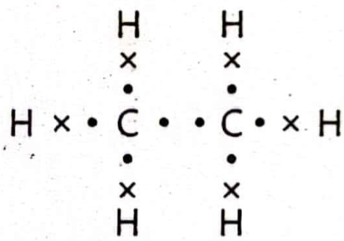
Methane



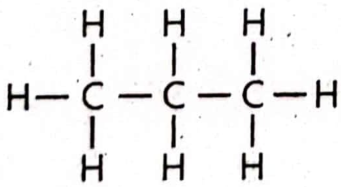


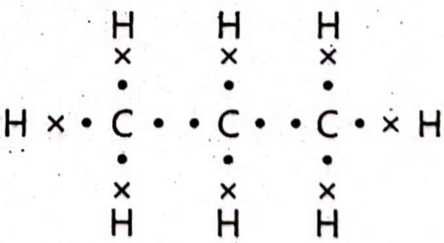
Ethane



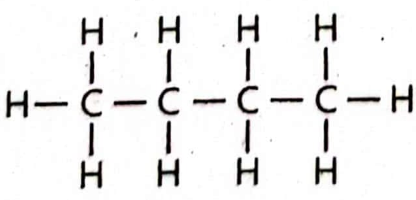


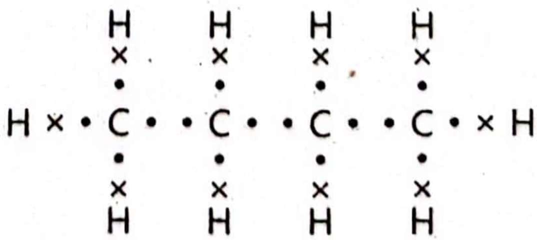
Propane





Butane





Exercise

How many methyl and methylene groups are present in each of the above compounds?

11.3 Preparation of Alkanes

Generally, any member of the alkane series can be prepared by the following methods.

1. Cracking of Higher Hyrocarbons

Cracking is the name of a process in which hydrocarbons with higher molecular masses (which are lesser in demand) are broken up into smaller hydrocarbons which are more in demand. This is done by heating the hydrocarbons at high temperature in the presence of a catalyst.

Fractional distillation of petroleum gives naphtha which consists of a mixture of liquid hydrocarbons. It is then heated at around 500°C in the presence of catalyst called zeolite to give hydrocarbons which have five to ten carbon atoms.

$$\underset{\text{(Mixture of hydrocarbons)}}{\text{Naphtha}} \xrightarrow[\text{Zeolite}]{\text{Heat}} \text{Alkanes and alkenes containing 5 to 10 carbon atoms}$$

Interesting Information!

Cracking of hydrocarbons helps balance the availability of petroleum fractions with the demand for them. When cracking transforms bigger hydrocarbons into small hydrocarbons, the fuel supply is increased. That helps to balance demand with supply.

2. Reduction of Alkenes and Alkynes

Alkanes can be prepared by reducing alkenes and alkynes with hydrogen gas in the presence of nickel metal as a catalyst. Methane cannot be prepared by this method. The reaction is also called hydrogenation of alkenes and alkynes and is an example of addition reaction. An addition reaction occurs when hydrogen (H_2) is added to an unsaturated compound.

$$\underset{\text{Ethene}}{CH_2 = CH_2} + H_2 \xrightarrow[200°C]{Ni} \underset{\text{Ethane}}{CH_3 - CH_3}$$

$$\underset{\text{Ethyne}}{CH \equiv CH} + 2H_2 \xrightarrow[200°C]{Ni} \underset{\text{Ethane}}{CH_3 - CH_3}$$

Interesting Information!

Addition of hydrogen to alkenes or alkynes is called reduction reaction. This reaction is used to prepare banaspati ghee and margarine.

3. Reduction of Alkyl Halides

Alkyl halides (R-X) can be reduced to alkanes with hydrogen generated by reaction of zinc metal with hydrochloric acid.

$$Zn + HCl \longrightarrow ZnCl2 + 2[H]$$

$$R - X \xrightarrow{[2H]} R - H + H - X$$

$$\underset{\text{Chloromethane}}{CH_3 - Cl} + Zn/HCl \xrightarrow{[2H]} \underset{\text{Methane}}{CH_3 - H} + H - Cl$$

11.4 Important Reactions

Alkanes are sometimes reffered to as paraffins which means "little affinity". This term describes their behaviour, for alkanes show little chemical affinity for other substances and are chemically inert to most laboratory reagents. However, alkanes do react with chlorine and oxygen under suitable conditions. The unreactivity of alkanes may be explained on the basis of the non-polarity of the bonds present in them. The electronegativity values of carbon (2.6) and hydrogen (2.1) do not differ appreciably and the bonding electrons between C – H and C – C bonds, are almost equally shared. This fact makes alkanes almost non-polar. In view of this, the reagents like acids, bases, oxidizing agents find no reaction site in alkane molecules to which they could be attached.

1. Halogenation

Alkanes give substitution reactions. The reactions which involve the replacement of hydrogen of alkanes by an atom or a group of atoms like halogen are called substitution reactions. Alkanes react with halogens especially chlorine to give alkyl halides. Since these substitution reactions are carried out in the presence of UV light, these are called photochemical substitution reactions.

$$\underset{\text{Methane}}{CH_4} + Cl_2 \xrightarrow{h\nu} \underset{\text{Chloromethane}}{CH_3 - Cl} + H - Cl$$

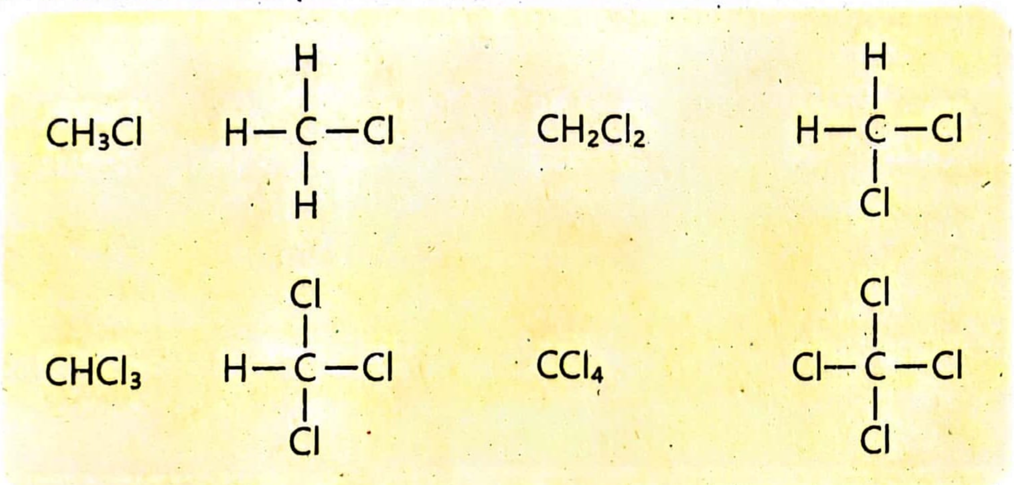
The reaction may proceed ahead and all the hydrogen atoms attached with carbon of the methane are successively replaced by chlorine atoms.

$$CH_3-Cl + Cl_2 \xrightarrow{h\nu} \underset{\text{Dichloromethane}}{CH_2Cl_2} + HCl$$

$$CH_2Cl_2 + Cl_2 \xrightarrow{h\nu} \underset{\text{Trichloromethane or chloroform}}{CHCl_3} + HCl$$

$$CHCl_3 + Cl_2 \xrightarrow{h\nu} \underset{\text{Tetrachloromethane or carbon tetrachloride}}{CCl_4} + HCl$$

Structural formulae of the products are shown below:



2. Combustion

Alkanes burn in oxygen or air to form CO_2 and H_2O with the evolution of large amount of heat. The reaction is called combustion.

$$\underset{\text{Methane}}{CH_4\,(g)} + 2O_2\,(g) \longrightarrow CO_2\,(g) + 2H_2O\,(g) + \text{heat}$$

$$\underset{\text{Ethane}}{2CH_3-CH_3\,(g)} + 7O_2\,(g) \longrightarrow 4CO_2\,(g) + 6H_2O\,(g) + \text{heat}$$

Interesting Information!

A mixture of natural gas (methane) and air may explode when ignited. This is the main cause of explosion at homes where gas leakage occurs.

Exercise

(1) In the reduction of alkyl halides with Zn / HCl, alkyl halide is being reduced. Which species in this reaction is being oxidized?

(2) During the combustion reaction of ethane, which bonds are being broken and which are being formed?

(3) What products other than CH_3Cl are formed when methane reacts with chlorine gas?

Key Points

1. The compounds obtained from plants and animals were named as organic compounds.
2. Organic chemistry is the chemistry of carbon compounds. The element carbon is unique in its behaviour.
3. Organic compounds are covalent in nature. Their melting points and boiling points are generally low.
4. Organic compounds containing carbon and hydrogen are called hydrocarbons. They are classified into saturated and unsaturated hydrocarbons.
5. Hydrocarbons containing a single bond between carbon and carbon and carbon and hydrogen are called saturated hydrocarbons or alkanes.
6. Alkanes can be named by a systematic way of nomenclature called IUPAC system of nomenclature.
7. Alkanes can be prepared by a number of different methods of preparation.
8. Although alkanes show least reactivity towards other compounds, they react with halogens and undergo combustion reactions.

Exercise

1. **Tick (✓) the correct answer.**

(i) Which other atom is almost always present along with carbon atom in all organic compounds?

(a) Oxygen (b) Nitrogen
(c) Hydrogen (d) Halogen

(ii) Which other metal can be used to reduce alkyl halides?

(a) Al (b) Mg (c) Ni (d) Co

(iii) If naphtha undergoes a combustion reaction what products do you expect it to form?

(a) Alkanes (b) Alkenes
(c) CO_2 and H_2O (d) Both alkanes and alkenes

(iv) Why does a mixture of zinc and hydrochloric acid acts as a reducing agent?

(a) Because zinc acts as a reducing agent.
(b) Because atomic hydrogen is produced with Zn / HCl which acts as a reducing agent.
(c) Because molecular hydrogen is produced with Zn / HCl which acts as a reducing agent.
(d) Because chloride ions are produced with Zn / HCl which act as a reducing agent.

(v) Which alkane will evolve the most amount of heat when it is burnt with oxygen?

(a) Ethane (b) Propane
(c) n-Butane (d) iso-Butane

(vi) Which reaction is not given by alkanes?

(a) Substitution (b) Combustion
(c) Addition (d) Cracking

(vii) Which hydrocarbon is responsible for explosions in coal mines?

a) Butane b) Pentane c) Methane d) Ethene

(viii) Which product will be formed when ethyl bromide (CH_3CH_2Br) is treated with Zn/HCl

(a) CH_4 (b) CH_3-CH_3
(c) $CH_3-CH_2-CH_2-CH_3$ (d) $CH_3-CH_2-CH_3$

(ix) Which of the following is not a process of halogenation of alkanes?
(a) Cracking (b) Chlorination
(c) Bromination (d) Iodination

(x) How many moles of oxygen will be required to completely burn propane?
(a) 4 moles (b) 5 moles (c) 3 moles (d) 6 moles

2. Questions for Short Answers

i. Differentiate between an organic and an inorganic compound.
ii. Why are organic compounds found in large numbers?
iii. Name the products which are obtained when natural gas is oxidized under controlled conditions.
iv. How naphtha fraction is decomposed to give lower hydrocarbons?
v. Write down the molecular formula, structural formula and the condensed structural formula for iso-butane.
vi. How are organic compounds useful for us?
vii. Write down the names of five such organic compounds which exist naturally?
viii. Give IUPAC name to the following compound.

$$\begin{array}{c} CH_3 - \underset{\displaystyle CH_3}{\underset{|}{CH}} - CH_2 - CH_3 \end{array}$$

ix. How do the melting and boiling points of alkanes change when we move from lower members to higher members?

3. Constructed Response Questions

i. Why do alkanes show little reactivity towards the other reagents?
ii. Why does a mixture of natural gas and air explode?
iii. How do you compare the melting and boiling points of inorganic and organic compounds?
iv. Reactions of alkanes with chlorine takes place in the presence of sunlight. What is the role of sunlight in the reaction?
v. How do you compare the boiling point of n – butane with that of iso – butane?
vi. Why are organic compounds not generally soluble in water?

4. Descriptive Questions

i. Describe the importance of organic compounds in daily life.

ii. Why is carbon so important as an element that the whole branch of chemistry is based on it?

iii. A carbon–carbon single bond (C – C) in alkanes does not behave as a functional group but a carbon carbon double bond (C = C) in alkenes does. Explain.

iv. Explain IUPAC system of nomenclature for alkanes.

v. How combustion reaction of alkanes is useful for us?

5. Investigative Questions

i. When natural gas valve is kept open in the kitchen, the gas spreads through the whole kitchen. This may cause an explosion. What is the reason of this explosion and how can you avoid it?

ii. "Neem" is a common tree grown throughout our country. Comment on the medicinal benefits of this tree.

iii. Name a few popular medicines which are, in fact, organic compounds?

Chapter 12

Empirical Data Collection and Analysis

Student Learning Outcomes

After studying this chapter, students will be able to:

- Explain that units are standardized for better communication and collaboration. (Some examples may include: In the field of chemistry, the International System of Units (SI) is used to measure physical quantities such as mass, volume, and temperature. This standardized system ensures that chemists worldwide can use the same units to measure and communicate their results, facilitating communication and collaboration in the field. Without standardized units, it would be difficult for chemists to compare their results with one another, and it would be challenging to develop consistent and accurate scientific models. For example, imagine if one chemist measured the mass of a substance in grams while another used ounces. The two measurements would be difficult to compare and combine, potentially leading to inaccurate or inconsistent results.)
- Identify SI units for abstract and physical quantities (some examples include mass, time and amount of matter
- Apply the concept that units can be combined with terms for magnitude, especially kilo, deci, and milli.
- Justify why chemists use cm^3, g and s as more practical units when working with small amounts in lab.
- Explain with examples how different tools and techniques can be used to manage accuracy and precision for inherent errors that arise during measurement.

Introduction

Science is a systematic study of this world through observation and experimentation. It is a method through which we make sense of this world in which we live.

Scientific research is done in all the countries of the world. But the way it is done is not identical everywhere. In order to make sure things are done properly and carefully, we need to share ideas and standardize our approach towards solving the problems.

One of the most usual problem which is faced by the scientific community is the issue of units. If scientists in one country are measuring lengths in metres and in another country in feet, then we will have to face problems in converting them. Comparing quantities in different units is not only confusing but the wastage of time as well.

For the reasons mentioned above scientists have agreed to adopt standard and user-friendly units called SI or System International Units. Things become a lot easier when we use these units.

The adoption of SI units is important in all branches of science because it makes communication easy worldwide. It allows scientists to share data easily.

SI units are preferred because they reduce the number of conversions needed to coordinate information among the scientists.

SI units use base 10, just like our number system. So, it is much easier to learn, remember and convert these units. These units are based on definite and precise standards. SI units are interrelated in such a way that one unit is derived from other units without conversion factors.

Interesting Information!

The following systems of units are commonly used in the world:
1. International System of Units (SI System)
2. Centimetre-gram-second System of Units (CGS System)
3. Metre-kilogram-second System of Units (MKS System)
4. Imperial System of Units

SI units are used almost everywhere in the world. It allows scientists to use a single standard in exchanging scientific data. This fact brings accuracy, consistency and universal understanding in scientific communication. A measurement taken in one part of the world can be easily understood and verified in another part without any confusion.

When scientists belonging to different countries and cultures collaborate on research, they need a common language to share their results. Using SI units enables scientists to compare results, replicate experiments and take benefit of each other work.

In conclusion we can say that SI units allow scientists to work together effectively, advancing the frontiers of our knowledge. All of this ensures safety, reliability, reproducibility and progress.

Exercise

1. What is the difference between reliable and reproducible results?
2. How SI units have brought harmony in the scientific community?

12.1 SI Units in Chemistry

There are seven base units in SI system for physical quantities out of which we use five in Chemistry. These physical quantities are length, time, amount of substance, mass and temperature.

12.1.1 Metre

It is the standard unit of length. Symbol **m** is used for metre. Metre is the distance travelled by light in vacuum in about 300 millionth of a second.

Fig 12.1: Metre rod

12.1.2 Kilogram

Its symbol is **kg** and it is the standard unit of mass. A block is kept in France which is taken as a standard unit of mass. It is also defined as the mass of 1000 cm^3 of water.

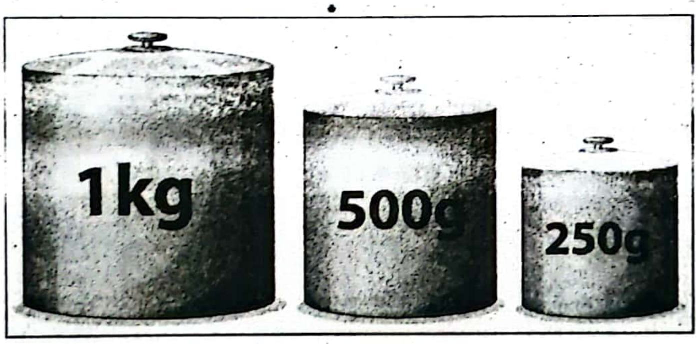


Fig 12.2: Different masses

12.1.3 Second

It is the standard unit of time with a symbol **s**. It is the time that elapses during 9,192,631,770 cycles of the radiation produced by the transition between two levels of the cesium-133 atom.

Interesting Information about SI Units

Symbols are not changed in plural forms. For example, 100 millimetres is written as 100 mm and not as 100 mms.

12.1.4 Kelvin

It is represented by **K** and it is the standard unit of temperature. It is $\frac{1}{273^{rd}}$ of the thermodynamic temperature of the triple point of water. It is a point at which all the states of water exist at the same time.

12.1.5 Mole

It is the base unit of the amount of pure substance and it is denoted by **mol**. It is defined as having exactly 6.022×10^{23} particles of a substance.

Table (12.1). Base Units of SI system

Quantity	Unit
Length	metre (m)
Time	second (s)
Amount of substance	mole (mol)
Mass	kilogram (kg)
Temperature	kelvin (K)

Derived Units

Apart from these base units, there are other quantities that are mathematically derived from these base units. Examples of the derived units used in chemistry are given in the following Table (12.2).

Table (12.2) Derived Units

Quantity	Unit
Volume	cubic metre (m^3)
Density	kg per cubic metre (kgm^{-3})
Area	square metre (m^2)

In addition to derived units, there are other specific quantities commonly used in chemistry Table (12.3).

Table (12.3) Specific Quantities used in chemistry

Quantity	Unit
Force	newton/N (kgm^{2-})
Pressure	pascal/Pa (Nm^{2-})
Energy	joule/J (Nm)

Since the SI system of units is a metric system, it is based around the number 10 for convenience. A set unit of prefixes has been developed which indicates whether the unit is a multiple or a fraction of the base ten. It allows the reduction of zeros of a very small number or a very large number. These SI prefixes also have a set of symbols that precede the unit symbol, Table (12.4).

Table (12.4) Prefixes used with SI units

Symbol	Prefix	
Mega	M	10^6
Kilo	k	10^3
Hecto	h	10^2
Deca	da	10^1
Deci	d	10^{-1}
Centi	c	10^{-2}
Milli	m	10^{-3}
Micro	u	10^{-6}
Nano	n	10^{-9}
Pico	p	10^{-12}

In Chemistry, we measure the masses of the reactants in grams. It is essential because the unit of measurement of molar mass consists of grams per mole. Therefore, given a mass measured in grams as well as a corresponding molar mass, enables us to find the mole of a substance. Moreover, in Chemistry the quantities involved in the laboratory are likely to be small. The choice of gram rather than kg is therefore sensible and normal. Using grams provides more manageable numbers for calculation and prevents the need for excessively large or small values.

Similarly, Celsius scale is most often used to measure temperature in Chemistry rather than Kelvin because it is more convenient to use it. Celsius scale has 100 divisions in total which makes it more compatible with the base ten format of SI system. Another reason is that it is easier to convert temperature on Celsius scale into Kelvin scale. The following equation is used for this conversion.

$$K = °C + 273$$

The unit of measurement of volume in Chemistry is cubic centimetre (cm^3) instead of cubic metre (m^3) because it is easy to measure and calculate with it and it is precise. In laboratory, we usually measure smaller volumes of liquid which are more manageable in cubic centimetre rather than cubic metre.

12.2 Tools and Techniques to Manage Accuracy and Precision

Measurement is the foundation for all experiments in science. Every measurement carries a level of uncertainty which is known as **error**. An error may be defined as the difference between the measured value and the actual value. For example, if two students use the same tool or instrument for measurement, it is not necessary that both of them get similar results. The difference between the measurements is called an error. An error may occur due to two factors: the limitation of the measuring instrument and the skill of the student making the measurement.

When we use tools meant for measurement, we assume that they give correct results. However, these tools may not always be right. In fact, they have errors that naturally occur and these errors are called **systematic errors**. Systematic errors may be removed by adding or subtracting a constant adjustment given to each measurement. Systematic error affect the accuracy of the measurement. All measuring instruments contribute to systematic error e.g. pipette, burette and measuring cylinder may deliver the volume slightly different from the one indicated by their graduation.

Another type of error which a student commits during measurement is called a **random error**. Random error causes one measurement to differ slightly from the next measurement. It comes from unpredictable changes during an experiment. The main reasons for random errors are limitations of instruments, environmental factors and slight variation in procedure. For example, when taking a volume reading from a measuring cylinder, you may read the volume from a different angle each time. Measuring the mass of a sample on a balance may give you different values as the surrounding air affects the balance. A random error often determines the precision of the experiment. The goal of any experiment is to obtain accurate and precise results.

12.3 Accuracy and Precision

Accuracy and precision are both ways to measure the correctness of results. They are used interchangeable in everyday life.

Accuracy measures how close results are to the true or known value. For example, the volume of a liquid is 26 cm^3 . A students measures its volume three times and find the result as 27cm^3. The student is not accurate because he has not calculated the exact result.

The closeness of two or more measurements to each other is called **precision**. For example, if you weigh a given substance five times and every time you get 3.2kg reading, then your measurement is precise but not necessarily accurate.

Precision is independent of accuracy. A student may be accurate but not precise and vice versa.

The exact mass of an object is 20 g. A student measures it and takes three readings as 17.3, 17.4 and 17.2. The student is considered as precise but not accurate. Similarly, another student measures the mass of the same object and gets readings as 19.8, 20.5 and 19.6. The second student is the more accurate but not precise.

Exercise

1. A student weighs a given substance three times, and each time he gets the reading 5.2g. The true weight of the substance is, however, 5.0g. Is the work done by the student **(i)** precise and accurate **(ii)** accurate but not precise **(iii)** precise but not accurate?
2. How will you avoid systematic and random errors?

Key Points

1. The subject of chemistry needs a consistent way to measure and to communicate the quantities like mass, volume, temperature, amount and time. To make sure that all of us can understand each other, scientists all over the world have adopted a common system of units which is based upon the metric system and it is called SI units.
2. There are seven base units and twenty two derived units in SI system but all these units are not used in Chemistry. In Chemistry we generally use five base units and three derived units.

Exercise

1. Tick (✓) the correct answer.

(i) Which of the following pairs of quantities may be measured in the same unit?
(a) Heat and temperature (b) Temperature and area
(c) Heat and work (d) Length and work

(ii) In which unit we usually measure the energy present in the food?
(a) Kilojoules (b Megajoules
(c) Calorie (d) Joule

(iii) What prefix is used for 10^{-12}?
(a) Mega (b) Pico
(c) Giga (d) Tessa

(iv) In SI unit of pressure is expressed in:
(a) Newton per metre (b) Newton per metre square
(c) Joule (d) Pascal

(v) Which symbol is used for kilogram in SI units?
(a) K (b) k
(c) Kgm (d) kg

(vi) What does a mole represent?
(a) Number (b) Mass
(c) Volume (d) Length

(vii) Which unit of volume should usually be used in Chemistry?
(a) Millilitre (b) Litre
(c) Cubic centimetre (d) Cubic metre

(viii) Express 0.000840 in scientific notation:
(a) 8.40×10^{-3} (b) 840×10^{-6}
(c) 8.40×10^{-4} (d) 84.0×10^{-5}

(ix) In SI units prefix nano means:
(a) 10^{-9} (b) 10^{-8}
(c) 10^{-11} (d) 10^{-12}

(x) $65^{\circ}C$ is equivalent to:
(a) 208 K (b) 338 K
(c) 403 K (d) 300 K

2. Questions for Short Answers

i. What is consistency of results?

ii. Why SI units are user friendly?

iii. Define systematic error and random error.

iv. What is reason behind a random error?

v. Does systematic error affect the accuracy?

vi. Which other systems of measurements are used apart from SI units?

vii. Define metre.

viii. Mention two benefits scientists get by using SI units.

3. Constructed Response Questions

i. Compare the units in SI system with those in MKS system?

ii. What are five basic SI units which are used in Chemistry?

iii. Explain the three units derived for the basic SI units.

iv. Explain why do we prefer to use smaller units of mass and volume in Chemistry?

v. What difficulties we expect to encounter if we use different units of measurement in daily life.

4. Descriptive Questions

i. What are our indigenous units of measurement of mass, volume and length?

ii. Elaborate the difference between precision and accuracy.

iii. How can you avoid systematic errors in your measurements?

iv. How do taking measurements in SI units ensure safety and reliability?

v. Can a student be both inaccurate and imprecise in his measurements?

5. Investigative Question

i. Elaborate the importance of using SI units in space exploration.

Chapter 13

Laboratory and Practical Skills

Student Learning Outcomes

After studying this chapter, students will be able to:

- Explain, with examples, the types of chemical hazards in the lab and suggest safety precautions. (Types of chemical hazards to be identified: flammable or explosive hazards, corrosive hazards, toxic hazards, reactive hazards, radiation hazards and asphyxiation hazards)
- Recognize the meaning of different chemical hazard signs in the lab and on chemicals.
- Recognize the importance of personal protective equipment (PPE) by correctly identifying the types of PPE needed for different lab. activities.
- Locate the nearest fire extinguisher and emergency shower.
- Show awareness of emergency procedures in the event of an emergency in the lab.

Introduction

A chemistry laboratory is a place where a student is trained to observe the physical and chemical characteristics of substances by following definite procedures. Before starting the laboratory work, a student should get himself familiarized with the layout of the laboratory and various fittings provided in the laboratory table as well as the side shelves. Students are expected to conduct themselves in a responsible manner at all times in the lab. They are advised not to work alone in the lab. Experiments should be performed in the presence of lab. instructor and other laboratory staff. All equipments should be checked before use whether they are working properly according to the requirements of the experiments. Determine the potential hazards related to any equipment or the experiment before beginning any work. Appropriate safety precautions must be observed at all cost. There must not be any crowding in the lab and students should stick to their work places at a safe distance from each other. Don't bring any food items in the lab. Never taste or smell any compound or a gas. If it is necessary to smell a gas it is always advised to waft the fumes or vapours towards your nose.

Warning signs are displayed when unusual hazards, hazardous materials, hazardous equipment or special conditions are expected. Do not pour chemicals down the drains and do not utilize the sewer for chemical waste disposal. Keep all sink traps and floor drains clean. Laboratory chemical waste can be disposed off in sewer or trash bin if they are non-hazardous materials. Acids and bases are first neutralized followed by sewer disposal. Hazardous waste material is transported to hazardous waste disposal site.

Misuse and mishandling of chemicals may create serious problems for the laboratory workers. A laboratory worker must use the chemicals according to the standard procedures keeping in view the particular hazards and precautions required for the safe use. The chemicals which can create problems for the safety of workers are cleaning agents, disinfectants, solvents, paints, compressed gas cylinders, mineral acids, carcinogenic chemicals etc.

Recognizing hazards which are commonly encountered in the laboratory helps to identify and minimize many of the health and safety problems. Most hazards which we might face while working in the laboratory fall into the following categories.

13.1 Chemical Hazards in the Laboratory

13.1.1 Flammable and Explosive Chemical Hazards

To start working in the laboratory requires great care, responsible behaviour and good attention. It is important to exercise extreme caution while working with delicate instruments, hazardous chemicals and open flames. If flammable and explosive chemicals are not handled in a safe and compliant manner, they can cause acute health problems. These problems may include burns, eye injuries, lung disease and suffocation.

Chemicals that cause a sudden release of pressure, gas and heat when they experience sudden shock are called explosive chemicals. Examples of chemicals which are expected to explode are picric acid, 2,4 -di-nitrophenyl hydrazine, benzoyl peroxide, nitrocellulose etc.

Flammable chemicals or mixtures are those which have a flashpoint around room temperature. Examples of flammable compounds are ethers, methylated spirit, benzene, acetone, petrol etc.

If you ever come across any chemical that you suspect to explode, do not attempt to move the container to avoid shock. Explosives can cause damage to people, windows, tables etc.

Avoid using a chemical that is hazardous. Look for its alternative. If you must use a potentially dangerous chemical, you must follow the underlying safety instructions.

1. Obtain prior approval from your teacher.
2. Always use smallest quantity of the chemicals.
3. Always conduct experiment in fume hood.
4. Remove all other chemicals and apparatus around you.
5. Inform other people working with you.
6. Always wear safety spectacles, gloves and lab. coat.
7. Always keep flammable compounds away from heat source.
8. Pour the flammable liquid very carefully.
9. Properly dispose off any hazardous waste.
10. Do not store flammable liquid in refrigerator.

13.1.2 Corrosive Hazards

Corrosive chemicals attack living tissues when they come in contact with them. They can be in the form of solids, liquids or gases. Such chemicals attack skin, eyes and respiratory tract and in the intestine as well. Whenever you work with corrosive chemicals, wear splash goggles instead of safety glasses and use a face shield.

Safety Precautions

1. Corrosive chemicals must be used in a fume cupboard to avoid breathing corrosive vapours.
2. While mixing concentrated acids with water, always add acid slowly to water and not vice versa.
3. Ensure eye wash and emergency shower is available.
4. Wash the affected area with soap and water and seek medical attention in case of emergency.

Examples of corrosive chemicals are mineral acids including HF, caustic alkalies, acetic acid (glacial) etc.

13.1.3. Toxic Chemical Hazards

A toxic chemical is a poisonous material which is capable of causing serious health problems. Mercury, benzene, chlorine, pesticides, ammonia, hydrogen cyanide are some examples of toxic chemicals. The following safety instructions may be ensured in case you intend to work with toxic chemicals.

1. Wear gloves, masks or other protective devices.
2. Keep the work area well ventilated.
3. Keep the toxic chemicals in original container.
4. Do not work alone.
5. Wash your hands with soap and water after you finished.
6. Always work in fume hood because toxic vapours can be formed during an experiment.
7. Seek immediate medical aid if you think you may have exposed to poisonous substance.

13.1.4. Reactive Chemical Hazards

The reactivity of chemicals is vital for the production of many other chemicals, pharmaceuticals and food products which are in our daily use. When chemical reactions are not properly performed, they may cause fires, explosions as they may evolve dangerous gases. These reactions may result to an extreme damage to life and property. Examples of reactive chemicals are calcium hydride, Na, Li, azides, picric acid, $AlCl_3$, benzoyl peroxide etc.

Safety Instructions

1. Handle reactive chemicals with utmost care. Segregate these from other chemicals while storage.
2. Appropriate measures should be taken before performing reactions with reactive chemicals. Utilize shield and heavy gloves.
3. Minimize the quantity required for experiment.
4. Glass equipment must be shielded by wrapping with tape.
5. After use carefully dispose off every dangerous material.

13.1.5. Radiation Hazards

When a person is exposed to a high dose of radiation, it can damage the functioning of tissues and organs and can cause vomiting, radiation burns, hair loss and radiation syndrome.

Radioactive materials that emit alpha and beta particles inflict extreme damage when inhaled or injected. Gamma rays cause external injuries. Medical x-rays produce ionizing radiation which can affect living tissues.

Safety Instructions

1. Keep radioactive sources shielded.
2. Avoid prolonged exposures to the radiation.
3. Stay inside as walls and ceilings can protect you from radiation fall out.
4. Never operate equipment that produces radiation without sufficient training.
5. Wear protective clothing, wear face mask.
6. Avoid contact of the material with bare skin.
7. Monitor exposure to radiation using badges etc.

13.1.6 Asphyxiation Hazards

It is a type of hazard in which a gas or vapours can cause unconscience or death through suffocation.

A sufficient level of oxygen is essential for normal breathing. If this level falls, it can create very dangerous situation. The exposed person has no warning and cannot realize that oxygen level has become low. If the level of oxygen decreases a person can feel rapid breathing, rapid heart rate, nausea and convulsions.

Examples of chemical asphyxiants are hydrogen cyanide, carbon monoxide, nitrogen, argon, helium, methane and carbon dioxide etc.

Safety Instructions

1. Store and use asphyxiant chemicals in well-ventilated areas with plenty of air.
2. Wear a full lab. coat, wear glasses and standard gloves, long trousers and closed-toed shoes.
3. Dispose off the waste strictly according to the instructions.
4. If exposed to such chemicals wash the exposed part with running water and seek medical attention.
5. When such a chemical is inhaled, remove the patient from the contaminated area and call appropriately trained person.

Exercise

1. Why flammable liquids are not stored in refrigerator?
2. Can you wear contact lenses in the lab.?
3. Under which circumstances explosive chemicals are likely to explode?
4. How will you dispose off acid and alkali waste after the experiment is finished?

13.2 Hazard Signs

A chemistry laboratory is a strict area where rigorous rules must be practised to avoid a chance of a deadly accident. A dangerous situation may arise not only for the individuals working there but for the whole area. In a laboratory there are several hazardous materials, sensitive equipments and specified areas for specific tasks. Proper warning signs ought to be posted on these areas to ensure that every person entering there must understand and act accordingly to maintain laboratory safety.

Several signs and symbols are posted in different areas of the lab. and bottles containing hazardous chemicals. These signs indicate that specific precautions must be observed according to the requirement of the sign posted there. If you see such signs, you must be alert and take extra care to maintain safety in that area, Fig (13.1).

Fig (13.1): Different Hazard Signs

Exercise

1. What does warning sign "caution" convey the message?
2. Name some explosive chemicals.

13.3 Personal Protective Equipment (PPE) in the Laboratory

Personal protective equipment should be made available to students to face any emergency situation which may arise in the lab. They are also useful to reduce exposures to hazardous chemicals. Proper protective equipment include such items as lab coat, protective glasses, face shields, apron, boots and hearing protection.

13.4 Location of Fire Extinguisher

Chemical laboratories using such materials which are likely to catch fire during experiments must have a portable fire extinguisher. This equipment can quickly be used to control a small fire if it is applied by a student individually. For this purpose all students should be well aware of the location where this fire extinguisher is placed. A training session should be held to train all the students to know how to handle and apply this fire extinguisher to extinguish the fire properly without any panic or harm to anybody.

Similarly the facility of a shower should also be made available in the lab whose location and working must be told to everybody working in the lab. In case of fire or any other emergency students should know how to face that emergency situation.

Exercise

1. Should emergency drills be compulsory or optional?

13.5 Emergency Situation in the Lab.

Students should make themselves aware of the actions that need to be taken in case of an emergency in a laboratory or if a person is affected. For this purpose periodic drills should be held with compulsory participation. Students should not only been given lectures but involve them practically to handle the emergency situations. During drill firefighting and other equipments must be checked whether they are in proper working order or not.

The following points should be kept in mind to cope with the emergency situation.

1. Stay calm and do not panic.
2. Alert people in the area to evacuate.
3. In case of fire, close doors to confine fire. Use fire extinguisher to put down the fire.
4. In case of chemical emergency adopt safety procedures as mentioned in article 13.1.
5. Call and assist emergency staff.

Key Points

1. Working in the laboratory requires care and responsible behaviour. Hazardous chemicals and open flames if not handled properly can cause health problems.
2. Chemicals can suddenly explode due to shock and heat. They require handling them with care.
3. Safety instructions should be followed strictly to avoid any damage due to flammable and explosive chemicals.
4. Corrosive chemicals affect skin, eyes and respiratory tract. To avoid such health problems corrosive chemicals must be handled in fume cupboard.
5. Chemicals are poisonous and cause great harm if not handled according to instructions.
6. Some chemicals are so reactive that they require special care in the laboratory.
7. Radioactive materials can affect living tissues and organs and cause other health problems. It is important to avoid longer exposure to radiation to stay healthy.
8. Asphyxiant chemicals are extremely lethal because they can cause suffocation. They must be used in well-ventilated places with protected dress.
9. Signs and symbols should be posted in the lab. and chemical bottles to let the people know their hazardous nature
10. Personal protection equipment are mandatory before you enter the lab.
11. Students shall know the location and operation of fire extinguisher and shower in the lab.

Exercise

1. Tick (✓) the correct answer.

(i) Safety in the chemistry laboratory is:

(a) the responsibility of the students only
(b) the responsibility of the professor only
(c) the responsibility of the lab. incharge only
(d) a shared responsibility

(ii) Accidents often result from:

(a) making mistakes
(b) failure to use common sense
(c) failure to follow instructions
(d) all of the above

(iii) The label "Warning" on a chemical bottle signifies:

(a) that the chemical can cause less serious injury
(b) that the chemical can cause serious injury
(c) that user should be careful when using chemical
(d) that user should open it only in the presence of a teacher

(iv) The label "Corrosive" on a chemical bottle indicates:

(a) that the material is an oxidizing agent
(b) that the material can degrade rapidly upon exposure
(c) that its contact destroys living tissue
(d) that the chemical can explode

(v) Example of highly toxic chemical:

(a) Ethanol (b) Acetic acid
(c) Potassium cyanide (d) Potassium permanganate

(vi) Example of self-reactive chemical:

(a) Potassium (b) Phenol
(c) Picric acid (d) n-Hexane

(vii) When diluting an acid with water:

(a) do it quickly

(b) do not stir the container

(c) always add acid to water

(d) always add water to acid

(viii) What should you do in case of a fire drill in the lab.?

(a) run to safety shower

(b) climb into the fume cupboard

(c) close gas valves and turn off all equipments

(d) carry chemicals out of the lab.

2. Questions for Short Answers

i. Name some corrosive chemicals.

ii. What type of safety precautions are adopted to avoid damage due to explosive chemicals?

iii. What type of damages can reactive chemicals cause?

iv. Indicate two such safety instructions which are required to avoid radiation.

v. Which chemicals can cause suffocation?

vi. Why signs and symbols are posted on lab. and chemical bottles?

vii. How fire caused by chemicals should be handled?

viii. Why emergency drills are important to face emergency situations?

3. Constructed Response Questions

i. How will you handle an emergency situation caused by fire due to short circuiting?

ii. What type of reactions should be carried out in fume cupboard?

iii. Put forward at least two suggestions to improve safety in the lab.

iv. Can you identify warning symbols posted for radiation and asphyxiant chemicals?

v. Why sudden shock can cause some chemicals to explode?

4. Descriptive Questions

i. Explain hazards due to explosive and toxic chemicals.

ii. Write down five such common safety instructions which are used to avoid all types of hazards.

iii. Explain the importance of warning signs and symbols to avoid any accident in the lab.

iv. Name some toxic chemicals. Describe the effects of spreading toxic gas in the lab.

v. A student has spilled over a corrosive and explosive chemical due to an accident. Which emergency measures you will take to tackle the situation.

5. Investigative Question

i. A few decades ago, a tanker carrying poisonous chlorine gas leaked and the gas spread over a large area in Lahore. The accident killed a few persons as well as animals. Give some concrete proposals to avoid such an accident in future.

Glossary

$2n^2$ Formula: Formula used for filling the electrons in shells.

Accuracy: It refers to how close a measurement is to the true value.

Acid Rain: Pollutant gases mixed with rain water come down as acid rain.

Activation Energy: Energy absorbed by the reactants and product molecules in order to be converted into the transition state.

Aerobic Respiration: The process of respiration in the presence of oxygen.

Aldehydes: organic compounds containing the aldehydic functional group

$$(-\overset{\overset{\displaystyle O}{\|}}{C}-H).$$

Aliphatic Compound: (Acyclic compounds) Compounds containing open chain of carbon atoms.

Alkanes or Saturated Hydrocarbons: Hydrocarbons in which all the four valencies are fully utilized or they contain single bonds only.

Alkenes: Compounds containing a double bond between two carbons atoms.

Alkyl Amine: Organic compounds containing amino group ($-NH_2$) as a functional group.

Alkyl Halides: A family of organic compounds containing halogen atom as a functional group.

Alkyl Radical: When an alkyl molecule drops one of its many hydrogen atoms.

Alkynes: Compounds containing a triple bond between two carbons atoms.

Anaerobic Respiration: Respiration without oxygen.

Aromatic Compounds: Compounds containing at least one benzene ring.

Arrhenius Acid: A chemical which gives proton (H^+) in water.

Arrhenius Base: A chemical compound which gives hydroxide ion (O^-H) in water.

Astrochemistry: It is the study of molecules and ions occurring in stars and interstellar space.

Atomic Mass: The mass of an atom of an element relative to the unit mass, which is 1/12th of the mass of C-12.

Atomic Number: The number of positively charged particles (protons) present in the nucleus of an atom.

Atomic Size: Average distance between the nucleus of an atom and its outermost electronic shell. Its units are m or pm.

Avogadro's Number: The huge number 6.022×10^{23} is called Avogadro's number.

Biochemistry: Which deals with the study of chemical compounds present in the living things.

Bronsted Acid: A compound which can donate proton.

Bronsted Base: A compound which can accept proton.

Carboxylic Acids: Organic compounds containing carboxyl group

$$(-\overset{\overset{\displaystyle O}{\|}}{C}-OH)$$ as a functional group.

Catalytic Converter: It is a device used in the exhaust of an automobile which converts harmful gases produced in the engine.

Catenation: The self-linking property of carbon.

Chemistry: It deals with the composition and changes in matter and the laws which govern these changes.

Colloidal Solution: A solution in which solute particles are bigger than those present in a true solution and which cannot be filtered.

Combustion: Burning of an organic compound in an excess of oxygen.

Concentrated Solution: A solution that contains a relatively large amount of a dissolved solute.

Concentration of a Solution: The amount of a solute which has been dissolved in a particular amount of solvent.

Coordinate Covalent Bond: When the shared pair of electrons is provided by one of the bonded atoms, a coordinate covalent bond is formed.

Covalent Bond: It is the force of attraction that arises between two atoms due to mutual sharing of an electron pair.

Crystal Lattice: Three-dimensional arrangement of ions.

Cyclic or Ring Compounds: Compounds in which carbon atoms are linked together to give a ring.

Dilute solution: A solution that contains a relatively small amount of dissolved solute.

Discharge Tube: A glass tube containing a gas at a very low pressure and provided with electrodes to study the passage of electricity through the gas.

Dynamic Chemical Equilibrium: When the rate of forward reaction becomes equal to the rate of reverse reaction, at this stage the reaction is said to be in a state of dynamic chemical equilibrium.

Electron: It is the smallest negatively charged particle present in all kinds of atoms. Its mass is 9.1095×10^{-31} kg and carries a charge -1.602×10^{-19} C.

Electronegativity: It is the power of an atom to attract the shared pair of electrons.

Empirical Formula: The formula of a compound, which shows the minimum ratio present between its atoms.

Endothermic Reactions: Those chemical reactions during which heat is absorbed.

Enthalpy of Reaction: Heat of reaction which takes place at constant pressure.

Enthalpy: It is the measurement of energy in a thermodynamic system.

Environmental Chemistry: In this branch, we study the chemicals and other pollutants present in the environment. It also covers the effects of these pollutants on living and non-living things.

Error: The difference between the measured value and the actual value.

Exothermic Reactions: Those chemical reactions during which heat is evolved.

Exotic states of Matter: These are not commonly encountered states of matter, for example, dark matter.

Extranuclear Portion: Area surrounding the nucleus of an atom.

First Ionization Energy: The minimum amount of energy required to remove an electron from the outermost electronic shell of an isolated gaseous atom. Its unit are kJ mol^{-1}.

Formula Mass: Formula mass is the mass of a compound relative to the unit mass which is 1/12th of the mass of C-12.

Functional Group: An atom or group of atoms or a bond whose presence imparts characteristic properties to the organic compounds.

Geochemistry: It covers the study of chemical composition of rocks and minerals.

Global Warming: The progressive warming of the Earth's surface due to the blanketing effect of CO_2, CH_4 and the water vapour present in the atmosphere.

Heat content: The total amount of heat energy present in a molecule under standard conditions.

Heat of neutralization: The heat given out during a neutralization reaction when one mole water is formed from an acid and base is called the heat of neutralization.

Heat of Reaction: Heat evolved or absorbed during a chemical reaction which takes place at any pressure.

Hydrated Salt: A salt with water molecular of crystallization.

Hydrocarbons: Compound of carbon and hydrogen only.

Inorganic Chemistry: The study of all other elements and their compounds except carbon and its compounds.

Inorganic Compounds: Compounds obtained form non-living things or mineral sources or synthesized in the laboratory by reacting metals with non-metals.

Intermolecular Forces: Forces between two separate molecules.

Ionic Bond: A bond formed due to the electrostatic force of attraction between oppositely charged ions.

Irreversible Reaction: Reaction which moves in one direction only; the reactants react to give the product.

Isomerism: The phenomenon shown by the organic compounds having the same molecular formula but different structural formula.

Isotopes: Atoms of an element having the same atomic number but different mass number.

Isotopic Abundance: The natural abundance of an isotope.

Ketones: Organic compounds containing the ketonic functional group

$$(R-\overset{\overset{O}{\|}}{C}-R).$$

Lipids: A group of organic compounds which serve as an energy reserve.

Liquid crystals: It is a state of matter whose properties are between those of liquids and solids.

Mass number: The total number of protons and neutrons present in the nucleus of an atom.

Medicinal Chemistry: Designing and synthesizing medicines or drugs which are useful to mankind.

Metallic Bond: When positively charged metal ions are held together by freely moving electrons, the bond formed is called a metallic bond.

Metallic Lustre: Shine present on metallic surfaces.

Modern Periodic Table: A table in which elements are arranged in ascending order of their atomic numbers.

Molar Mass: The mass of an element or a compound which contains Avogadro's number of particles.

Molecular Formula: The formula of an element or a compound which tells the actual number of atoms present in the molecule of that element or a compound.

Molecular Mass: Molecular mass is the mass of the molecule of an element or a compound relative to the unit mass, which is 1/12th of the mass of C-12.

Neutralization: Acids and bases react together to form salts and water and, in this way, they neutralize the properties of each other. This reaction is called neutralization reaction.

Neutron: It is the smallest neutral particle present in the nucleus of atoms. Its mass is slightly more than that of a proton.

Nuclear Chemistry: This branch deals with the reactions taking place in the nucleus of an atom.

Nucleus: Central part of an atom where most of its mass is concentrated. Its size is very small as compared to the size of the atom.

Octet Rule: When an atom has eight electrons in its outer most shell, it is said to be stable and does not combine with other atoms to reduce its energy. This is called octet rule.

Orbit: the circular path of an electron around the nucleus.

Organic Chemistry: The branch of chemistry in which we study the compounds of carbon.

Organic Compounds: Compounds obtained from living or plant and animal sources and which can be synthesized in the laboratory. All the organic compounds contain carbon as an essential element.

Oxidation: A process in which an electron or electrons are lost.

Oxidising Agent: A substance which accepts an electron or electrons.

Parts per Million: One part per million parts of the solid, liquid and gas mixture in which pollutant is formed.

Physical Chemistry: This branch investigates how substances behave on an atomic or molecular level and how physical laws govern the specific characteristics of atoms and molecules.

Plasma: It is the fourth state of matter. It is composed of particles with very high kinetic energy.

Pollutant: Any solid, liquid or gaseous substance present in such concentration as may be injurious to health.

Polymer Chemistry: It focuses on the properties, structure and synthesis of polymers and macromolecules.

Precision: It refers to how close measurements of the same item are to each other.

Proton: It is the smallest positively charged particle present in all kind of atoms. The mass of this particle is equal to the mass of hydrogen nucleus (H^+).

Radioactive rays: Rays emitted from radioactive elements or their compounds, which can cause fogging of the photographic plates.

Radioactive Isotopes: Isotopes of elements which throw out excess energy in the form of radiation.

Radiocarbon Dating: A method for finding out the age of a historical object containing organic material with the help of ${}^{14}_{6}C$.

Random Error: Error which a student commits during measurement.

Reducing Agent: A substance which loses an electron or electrons.

Reduction: A process in which an electron or electrons are gained.

Reversible Reaction: A chemical reaction, which takes place in both directions, forward as well as backward.

Saturated Solution: A solution, which contains the maximum amount of a solute at a particular temperature and which is unable to dissolve further amount of solute in it.

SI Units: A common system of units based on metric system.

Solubility: The amount of a solute in grams which has been dissolved in 100g of a solvent at a particular temperature to prepare a saturated solution.

Strong Acid: An acid which ionizes completely in water.

Strong Base: A base which can ionize completely in water giving excess of hydroxide ions.

Structural Formula: The formula which shows the arrangement of atoms in a compound.

Substitution Reaction: A reaction which occurs when an atom or a group of atoms from the reagent displaces an atom or group of atoms from the organic reactant.

Supercritical Fluids: They are highly compressed gases which show the properties of both gases and Liquids.

Surrounding: Everything else which does not fall in the system.

System: Anything under consideration for the purpose of study.

Systematic Error: Error which naturally occurs when we use tools for measurement.

Transition Elements: Elements having incomplete penultimate (next inner to the outermost) electronic shell.

Transition State: A state of molecules when they are undergoing breakage or formation of bonds.

Unified Atomic Mass Unit: Unit of a new scale, which is equal to 1/12th of the mass of C-12.

Unsaturated Hydrocarbons: Hydrocarbons containing double or triple bonds.

Unsaturated Solution: A solution, which can dissolve further amount of a solute at a particular temperature.

Vital Force: The imaginary supernatural force which was supposed to be present in all those compounds which were obtained from living things.

Water of Crystallization: The number of water molecules present in the crystals of a solid.

Weak Acid: An acid which ionizes partially in water.

Weak Base: A base which ionizes partially in water.